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FACTS AROUND US

C. LLOYD MORGAN.

FACTS AROUND US:

SIMPLE READINGS
IN
INORGANIC SCIENCE.

WITH EXPERIMENTS.

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INORGANIC SCIENCE;
WITH EXPERIMENTS.

BY

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TO THE TEACHER.

Two things I would ask of you, My Friend in Science, if you make use of this little book: First, to perform carefully and accurately every experiment therein described, and to see that each individual in your class quite understands what is used and what is produced. Secondly, to draw on the black-board outline sketches of all the apparatus you use, writing the names of the substances employed to the left of your figure, and the names of the substances produced to the right of your figure; and to make each individual in your class do the same in a note-book specially kept for that purpose, and for additional notes on the subjects of the lessons.

And two bits of advice I would offer: First, that, apart from *vivâ voce* work, you should at the beginning of each hour require written answers, from each individual in your class, to six or eight questions on the preceding lesson, these answers to be subsequently corrected by you, and returned with marks assigned. Secondly, that you should repeat the experiments from time to time, not fancying that, because your class has seen an experiment once, it will learn nothing by seeing it again, but rather believing that, until the experiment has been repeated, it is but half performed.

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ERRATA.

p. 32, line 6, *after the words* closed glass tube, *insert the words* containing mercury.

p. 121, line 10 from bottom, *for* nitric acids *read* nitric acid.

p. 124, line 8 from bottom, *for* remains *read* remain.

[illegible]

FACTS AROUND US:

SIMPLE READINGS
IN
INORGANIC SCIENCE.

WITH EXPERIMENTS.

out the candle and shaking up the lime-water in the vessel, the liquid will turn white and milky. This shows the presence of carbonic acid gas.

When the candle burns it disappears, and is apparently destroyed. But it is not really destroyed: that is to say, the matter of which it is composed is not destroyed. That matter has merely united with the oxygen of the air, and assumed new and invisible forms.

II. ANIMALS AND PLANTS.

Living creatures are of two kinds, Animals and Plants. At first sight it seems easy to say how these two kinds differ from each other; a horse, a whale, or a butterfly, for example, differ very much from an oak, a palm, or a fern. But there are some minute and lowly organized animals and plants, between which we find great difficulty in telling the difference. Perhaps the first distinction that would strike us is this, that animals can move about from place to place, while plants are fixed or rooted to one spot. But there are exceedingly minute plants, only to be seen under the microscope, which are quite as active as the minute animals of the same size. And, on the other hand, some sea-shells, such as the oyster, are as firmly rooted to one spot as plants. This distinction, therefore, which holds good for most animals and plants, does not hold good for all.

There is another distinction between most animals and most plants, with which we have now more especially to deal, namely, that animals breathe-in oxygen, and breathe-out carbonic acid gas; while plants, in daylight, and especially in sunshine, take in carbonic acid gas and give out oxygen. We, as human beings, belong to the animal kingdom: and we take in oxygen with every breath we inhale, and breathe out carbonic acid gas. In the last lesson we saw that the carbonic acid gas, formed when a candle burns, has the property of making lime-water turn milky, owing to the formation of minute solid particles of carbonate of lime. If now I breathe by means of a thin glass tube through clear lime-water, it rapidly becomes milky, the milkiness being caused by the carbonic acid gas which I breathe out. If we put a bunch of green leaves in a glass of water in the sunlight, we shall see that the leaves soon become beaded with little bubbles of oxygen gas. This is seen still more clearly in a small aquarium, in which water plants are growing. Here we may often see the bubbles rising quite rapidly to the surface. If we keep fish and other animals in an aquarium without any plants in the water, we must be careful to change the water frequently. For the fish soon use up all the oxygen that is dissolved in the water, and give up carbonic acid gas which also dissolves in the water. If, therefore, a fresh supply of water containing some more oxygen be not given them the creatures die. If, however, there be plants in the water it need not be changed so frequently. The plants absorb the carbonic acid gas which the animals give out, and they set free oxygen, some of which dissolves in the water, and makes it fit for the animals to live in. If an animal be kept in a confined space it will die of suffocation. We need not perform any experiment which entails suffering on a creature, but we can find in history a sad instance of the death of many human beings in this way. In 1756 Surajah Dowlah took Calcutta, and confined one hundred and forty-six English people in a dungeon known as the Black Hole of Calcutta. Many of them died of suffocation.

By the joint action of animals and plants the proportion of carbonic acid gas in the air is kept constant. Every animal that breathes produces carbonic acid gas, and this has been going on for thousands of years. And yet the atmosphere only contains about one cubic foot of carbonic acid gas for every 2000 cubic feet of air. Why? Because the plants absorb all the carbonic acid gas that animals breathe forth. And what do they do with all this carbonic acid gas? They take up out of it some of the stuff of which their wood is composed, the carbon, and set free the oxygen which before was combined with the carbon. The plants, therefore, absorb the carbonic acid gas, use the carbon for building their leaves and stems, and render back the oxygen sweet and fresh for the use of men and animals. It is only the higher plants, however, (those which have green leaves,) that do this even in sunshine; colourless plants and fungi take in oxygen, and give out carbonic acid gas, as do all plants in the dark.

Another distinction between animals and plants, is that animals must feed on animal or vegetable matter, while plants can feed on carbonic acid gas, and water, with some mineral matter. Plants, in fact, separate the carbon and oxygen in carbonic acid gas; the carbon goes into their leaves, the oxygen into the air. Then comes the animal, eats the plant and breathes the air, and so brings the two into contact that they unite again to form carbonic acid gas, which other plants will decompose. Even this distinction between animals and plants, in the matter of their food, does not hold good for some little creatures, which seem half their lives to eat like plants, and the other half to eat like animals.

III. CHALK.

COMMON chalk is a kind of limestone. It is found in many parts of Europe, and among others in the south of England, where the white chalk cliffs gave rise to the name, **Albion.** If water contain some carbonic acid gas dissolved in it, it can dissolve chalk; and the rivers which are fed by streams coming from the chalk downs contain a good deal of chalky matter in solution. This is carried out to sea. And near the surface of the ocean, in temperate and tropical latitudes, there live minute creatures which take the chalky matter out of the water, and with it build their shells. These shells are less than one hundredth of an inch in diameter, and are built up of chambers perforated with very minute holes. From this fact they are called **foraminifera** (Lat. *foramen*, a hole; *fero*, I bear). Through these holes the creatures can protrude the jelly-like substance of which they are composed. When these little creatures die, their shells sink down to the bottom; and they sink in such numbers that there is a constant rain of these little shells on to the bottom of some parts of the ocean. There they collect and form a greyish white mud called **ooze.** From the name of the most abundant shell it is often called,

globigerina ooze. As this process is constantly going on, the creatures dying day by day in thousands, and the little shells sinking down to the bottom, great quantities of ooze must thus collect at the bottom of the ocean. Now by examining chalk under the microscope it has been found that this substance is formed of a material very similar to ooze, in which the most abundant shell is the very same globigerina. And there can be very little doubt that chalk was once formed in the past in very much the same way as ooze is being formed at the bottom of the Atlantic to-day.

Chalk is composed of carbonate of lime, a substance made up of lime united with carbonic acid gas. We must notice that though this carbonic acid is a gas when it exists in the free state, it unites with the lime to form **solid** chalk. If chalk be thrown into a lime-kiln and heated, the carbonic acid gas is driven off by the heat, and the lime remains behind in the kiln. This lime, which is used for making mortar, is called **quicklime.** If cold water be poured on to the quicklime it becomes so hot that the water hisses on the hot lime, and steam is formed. This is because the quicklime and the water unite chemically, and chemical union, as we have already seen in the case of the burning hydro-carbon in the candle flame, very often gives rise to much heat. The lime is now called **slaked lime**: it is a dry powdery substance, for the water seems to have quite disappeared. Just as the carbonic acid gas united with lime forms a solid substance, so does the liquid water united with the lime form a solid substance. If now we shake up a little of this slaked-lime in water, some of it will dissolve, and we shall have **lime-water**. And we have already seen that if we shake up lime-water with carbonic acid gas the clear liquid becomes milky, owing to the formation of minute solid particles of carbonate of lime. In the lime-kiln the lime and the carbonic acid gas are forcibly separated by the application of heat: in the lime-water the lime and the carbonic acid gas come together again, and once more unite. We start with carbonate of lime, and we end with carbonate of lime. The carbonate of lime we start with is made up of the minute chalk-building globigerina shells: but the carbonate of lime we end with consists of minute particles thrown down or **precipitated** from the lime-water.

IV. THE MAMMOTH CAVE.

We have seen in the last lesson that there are many minute creatures in the sea that are busily employed in extracting the chalky matter dissolved in the water. Of this chalky matter their shells are formed. And when they die their shells sink to the bottom, and there, accumulating in vast numbers, form a grey mud called globigerina ooze, which is a kind of modern chalk. Besides these globigerinæ and other foraminifera, there are many other living creatures which extract the chalky matter out of the water (just as we extract the substance of our bones out of our meat and drink), and with it build their skeletons; which skeletons collecting in immense numbers form beds of limestone rock. Corals for instance do so: and there are immense masses of modern limestone which are entirely composed of dead coral. Coral reefs surround many of the Pacific Islands, and a huge reef of the same material runs for a thousand miles along the east coast of Australia.

Now it is found that very much of the dry land which forms our continents is made up of limestone of some sort; and nearly all this limestone has been produced, long ages ago, by the action of living creatures, such as the foraminifera, the corals, or others. But what we have now especially to notice about this limestone is, that it is very apt to be full of great caves and caverns. In Yorkshire, where there is much limestone, there are many great caves, such as the Clapham Cave; in the limestone in Derbyshire there is the celebrated Peak Cavern. In South Africa there are, in the Oudtshorn district, a series of caves, called the Cango Caves, which run underground for many miles, and have never been thoroughly explored. But one of the largest caves in the world is in Kentucky, in America. It is known as the Mammoth Cave, and is said to have been discovered by a hunter who, in 1809, followed a bear into what seemed a small hole in the earth. This hole, however, led into a large underground tunnel. This main tunnel of the Mammoth Cave, from which many minor passages open out, is an old deserted water-course, and runs, through a succession of great arches and domes, to a point about six miles from the entrance, where it is abruptly closed by fallen rocks. Its average width is about 60 feet, and its height 40 feet. About 190 feet below the level of the main cave there are two rivers, called the Styx and the Echo. Once the water which flows in this lower course, flowed in all probability through the main cave. It has been estimated that the united length of all the passages in the Mammoth Cave is from 150 to 200 miles. Some 12 million cubic yards of limestone have been removed by the waters. In the cavern there are many deep shafts which have been formed by subterranean waterfalls; their sides resemble curtains delicately carved in stone. In one place columns of white limestone springing from floor to roof have caused this part of the cave to be called the Gothic Chapel.

Let us now inquire how the Mammoth Cave has been eaten out of the solid rock, and how these crystalline columns have been formed. Pure water will not dissolve limestone. But if the water contain in solution some carbonic acid gas, it dissolves some of the limestone without difficulty. Now there is, as we have already learnt, some of this carbonic acid gas in the air; and whenever leaves and twigs of wood decay some of this gas is formed. The water which enters the Mammoth Cave contains some of this carbonic acid gas dissolved in it; it can thus dissolve some of the limestone rock. And in the course of ages the water has, by means of the carbonic acid gas invisibly dissolved in it, been enabled to eat out of the solid rock the great Mammoth Cave.

In many places water drips from the roof of the cave. This water has, by means of the carbonic acid gas it held in solution, dissolved some of the limestone. But when a water-drop exposes a large proportion of its surface to the air the carbonic acid gas escapes. Then the water, without the carbonic acid gas, can no longer hold dissolved the limestone it held before. Some of the limestone it deposits on the roof of the cavern, in the form of a beautiful stone-icicle, called a **stalactite** (Gr. *stalazein*, to drop or drip). Some it deposits on the floor of the cave in the form of a rounded dome. This is called a **stalagmite**. By the constant succession of drops, stalactite and stalagmite grow both in length and in thickness, until they meet and form a beautiful crystalline column of limestone, such as may be seen in the Gothic Chapel of the Mammoth Cave.

If we examine a piece of a broken stalactite, we shall see that it is composed of a series of crystalline rings of limestone. Very likely there may be a hole running through the middle of the specimen. If so we may be sure that some of the water constantly found its way down this hole. Miniature stalactites may often be seen hanging from the under surface of the arches of a bridge. The lime is in such cases obtained from the mortar.

V. CARBONIC ACID GAS.

When a candle burns in the air, carbonic acid gas is formed; and when chalk is heated in a lime-kiln, carbonic acid gas is given off. But there is a difference in the way in which the carbonic acid gas is produced in these two cases. When the candle burns the carbon of the hydro-carbon unites with the oxygen of the air; the carbon and oxygen are brought together from different sources. But when the chalk is heated the carbonic acid gas **already contained in the chalk** is separated from the quicklime. The former is a case of **chemical union**; the latter a case of **chemical decomposition**.

There is another way of decomposing the chalk beside heating it in a lime-kiln. If we put some pieces of common chalk in a glass flask, and just cover them with water, and then add some **hydrochloric acid**, the chalk will dissolve, and bubbles of gas will be seen rapidly to rise to the surface causing what is called effervescence. These bubbles of gas are the carbonic acid gas, hitherto locked up in the particles of the chalk, making their escape under the influence of the hydrochloric acid. If now we fit the flask with a cork, through which a long curved glass tube passes, we may lead off the gas from the flask, and make some experiments with it.

In the first place we may lead some of the gas into some water in which a few grains of blue **litmus**, a vegetable colouring matter, have been dissolved. We shall find that the blue gradually changes to red. This shows that some of the gas has dissolved in the water; and it also shows that the gas has acid properties, for all **acids turn blue litmus red.** If, for example, we add a single minute drop of the hydrochloric acid, with which we decomposed the chalk, to a solution of blue litmus, that solution will at once turn bright red. We may notice, in passing, that the red produced by the hydrochloric acid is rather brighter than that produced by the carbonic acid gas.

A second experiment is to lead some of the gas into some lime-water in a test tube. The lime-water turns milky, the milkiness being due to the formation of minute particles of carbonate of lime. But if more carbonic acid gas be led into it, the lime-water, thus made milky, gets quite clear again. This is because the additional carbonic acid gas enables the water to dissolve the particles of carbonate of lime. Here we have, in fact, an experimental illustration of the fact before mentioned, that water containing carbonic acid gas dissolves limestone (as in the Mammoth Cave) or chalk (as in the chalk downs), these substances being composed of carbonate of lime. Water with carbonate of lime thus dissolved in it is said to be 'hard' water. There are two ways of making it 'soft,' that is of getting rid of the carbonate of lime. The first is to boil the water. This forces the carbonic acid, which enables the water to dissolve the carbonate of lime, to escape. The carbonate of lime, therefore, separates from the water in the form of solid particles. These particles sometimes collect at the bottom of kettles, in which hard water is constantly boiled. It there forms quite a thick layer, sometimes called a 'fur.' A second way of getting rid of the hardness is to add more lime-water. The lime in the lime-water unites with the extra carbonic acid gas, and forms solid particles of carbonate of lime. But now the extra carbonic acid gas has been removed, the old carbonate of lime particles can no longer be held in solution. The new carbonate of lime particles, therefore, and the old carbonate of lime particles separate at the same time from the water, and both sink to the bottom together, leaving clear soft water above them.

A third experiment may be made as follows: We may place a glass vessel on one pan of a delicate pair of scales, and put weights on the other pan so as exactly to balance the glass vessel. If now we let some of the carbonic acid gas flow from our flask into this balanced vessel, we shall find that the gas makes the vessel heavier, so that the weights no longer balance it. It is just as if we had poured a little water into the vessel. This shows that carbonic acid gas is heavier than air. It is, in fact, about one and a half times as heavy as air. If, however, we leave for a short while the glass vessel into which we passed the carbonic acid gas, we shall find that it resumes its former weight. For though the carbonic acid gas is heavier than air, it will not remain in an open vessel, but will gradually escape into the air. This is because the particles of all gases try and get as far apart from each other as possible.

If we have not a pair of scales for this last experiment, we can still exemplify the fact that carbonic acid gas is heavier than air. We may put a little lime-water at the bottom of a test tube, and may then pass some of the gas into an empty bottle. After a few minutes we may pour the carbonic acid gas out of the bottle in which it has collected, into the test tube containing the lime-water. On closing the test tube with the thumb and shaking, the lime-water will become milky, showing that some of the carbonic acid gas has flowed downwards out of the bottle into the test tube. And the downward flow shows that the carbonic acid gas is heavier than air.

VI. WOOD AND COAL.

It is only under the influence of sunlight that green plants are able to decompose carbonic acid gas, taking up the carbon for their own use and setting free the oxygen for the use of men and animals. Without sunlight the higher plants could not grow. A noble tree has locked up in its fibres the virtue of many days of bright sunshine. Suppose the tree be felled and cut up into logs and billets for burning: then as the wood burns the carbon unites again with the oxygen from which it was separated, carbonic acid gas again being formed. And in the dark winter's night we get back some of the solar warmth and light which was locked up, as it were, in the fibres of the wood. If the tree grow old and die, the wood begins to rot and decay. In this case, too, the carbon and oxygen unite again to form carbonic acid gas; but here the union takes place very much more slowly. Now, suppose the wood, instead of either being burnt or decaying, gets buried in the earth, so that the oxygen of the air cannot get to it: then the carbon will be kept safe and preserved from destruction. Coal is woody matter that has been buried, and thus preserved in this way. The plants of which the coal is made up lived long ages ago, and have been buried beneath enormous masses of earth and rock. The great pressure of all this rock has squeezed the woody matter into a hard and compact mass, and this pressure and other causes have converted it into the hard, black, shining substance coal.

Great quantities of vegetable matter are now being buried in some parts of the world. On the western side of the Mississippi river in North America there is a vast "sunk country," caused by an earthquake which took place in 1811. Here there are many great lakes. And all round the edges of these swampy lakes, and out into the shallow water, cypress and other trees grow in luxuriant profusion. Here those trees live and die, their leaves and their fallen trunks floating for a while on the stagnant waters. Soon they get water-logged and sink to the bottom; and as other generations of trees die, their remains swell the amount of vegetable matter which is thus slowly but surely being deposited. And this vegetable matter is very pure, it is mixed with very little sand or mud. For all the water which finds its way into the swampy lake, has to filter through reeds and rushes, and a network of water-weeds. The water which reaches the lake is pure water so far as suspended matter is concerned, and thus a layer of pure vegetable matter is laid down on the clay on which the cypresses grow. The purity of the vegetable deposit is sometimes shown, during an unusually hot season, when a portion of one of these swamps becomes dried up. For if the wood is then accidentally set on fire, as is sometimes the case, pits are burnt into the ground many feet deep, or as far down as the fire can descend without meeting with water; so pure is the mass of vegetable matter accumulated in these swamps.

The plants and trees, however, of which the very ancient vegetable accumulation we call coal is composed, were very different to cypresses, and indeed to any trees that we are accustomed to see. Some of them were like gigantic club-mosses. Certain thick beds of coal seem to be entirely composed of very minute seeds or spores of such club-mosses. But whatever the trees were like, they lived, just as our trees do, by decomposing carbonic acid gas, and storing up the carbon. And they were able to do this only through the influence of sunlight. The carbon was then buried, and has remained hidden in the earth for long ages. At last man comes and digs out the remains of these old plants stored up as coal. He burns the coal and lets the carbon unite again with oxygen; and while they unite he feels, as it were, the pleasant warmth and light which long ages ago the plants drank in from the sun. Then the carbonic acid gas, formed by the union, goes forth into the air, and is, ere long, seized by living plants, by which the carbon and oxygen are again separated.

VII. CHARCOAL, BLACK-LEAD, AND DIAMOND.

COAL is not pure carbon. It contains besides carbon three gases, **oxygen, hydrogen,** and **nitrogen.** These gases, however, being in combination with each other and with the carbon, have lost their gaseous form. The amount of carbon varies from seven to nine parts out of every ten. Wood contains less carbon still: about five parts out of every ten being carbon. But if either coal or wood be strongly heated, the oxygen of the air being carefully excluded, the gases are driven off in combination with some of the carbon, and nearly pure carbon remains behind. The carbon made by heating wood is called **charcoal**; that made by heating coal is called **coke.** Both charcoal and coke are useful substances, and are therefore prepared on a large scale. But we must be content with making some experiments on a small scale. If we place half a dozen wooden matches, the prepared tips of which have been cut off, in a test tube, and, nearly closing the mouth of the test tube with a loosely-fitting cork, heat the test tube over a lamp, we shall see that the matches at first begin to blacken. The tube should be held with its mouth slightly inclined downwards. Heavy strongly-smelling white smoke is given off, the tube becomes blackened round the matches, while liquid drops of discoloured water run downwards towards the cork. Soon the blackened and shrunken matches will no longer give off any smoke; and we may then let the test tube cool and examine the pieces of charcoal, into which our matches have been converted. We shall find that the charcoal is soft, light, and readily soils the fingers. The extreme lightness is due to the fact that the charcoal is a spongy mass full of air. We may throw one of our pieces of match-charcoal on to water and it will at first float. But if we leave it in the water, the water will, after some time, fill up all the little air cavities, and the match will then sink to the bottom. To make an experiment with the coal the best way is to take a common clay tobacco pipe and fill the bowl about three quarters full of pounded coal. We must then cover over the top of the bowl with clay. After allowing the clay to harden we may put the bowl in the fire or heat it over a lamp. After a little while we shall see smoke issuing from the stem of the pipe. If we apply a match to this, we shall find that it burns with a bright flame. It is in fact coal-gas. When the coal is heated on a large scale this coal-gas is carefully purified, and can then be used to light our rooms and our streets. When no more gas is given off from our tobacco pipe, we may take it out of the fire or remove the lamp. After it has cooled we may break it open, and we shall then find a caked mass of coke inside the bowl. The coke is harder than the charcoal, and does not so readily soil the fingers. Some kinds of coal are better than others for making gas. There are indeed some kinds that produce hardly any gas; they are composed of nearly pure carbon. All the gas has been driven out of them by pressure, heat, &c., while they were still in the earth.

The black lead or graphite, of which pencils are made, and which is also used for polishing grates, is nearly pure carbon. It is a soft, black, shining substance found in some of the oldest rocks, especially in Cumberland. But no one quite knows how it was formed in these rocks.

Another and still purer form of carbon is the diamond. It is very different to graphite, for instead of being black and soft it is very often as transparent as glass, and is the hardest stone there is. All the diamonds that were used before 1728 came from Hindustan or the East Indies. In that year the mines of Brazil were discovered. It was not until 1868 that diamonds were discovered in South Africa. The weight of diamonds is given in carats, a carat being about four grains. The grain of corn was the smallest Arabic weight: four of these were equal to one sweet pea, called in Arabic carat. The 'Great Mogul' diamond, in the imperial sceptre of Russia, is estimated to weigh 779 carats, and to be worth nearly £5,000,000 sterling. Carbon, therefore, when pure or nearly pure, exists in three quite distinct forms; charcoal, graphite or black-lead, and diamond. If any one of these three forms be heated in pure oxygen gas it burns, charcoal very readily, graphite with difficulty, diamond less readily still. And in every case carbonic acid gas is produced.

Let us try and understand how men of science have been able to find out that three substances so different as charcoal, black-lead, and diamond, are all different forms of the substance we call carbon. What they do is this: They weigh out carefully 12 grains of charcoal, and completely burn it in oxygen. They collect all the carbonic acid gas produced and weigh it too. They find that 44 grains of carbonic acid gas are produced. They then weigh out 12 grains of black-lead and burn it: and they find that 44 grains of carbonic acid gas are produced in this case too. They then weigh out 12 grains of diamond and burn it: and they find that it too produces 44 grains of carbonic acid gas. In each case, then, 12 grains of the substance produce, when completely burnt, 44 grains of carbonic acid gas. Now remember that when the substances are burnt they are not really destroyed: they merely unite with the oxygen, and assume the new form of carbonic acid gas. The carbonic acid gas in each case contains all the substance burnt, for no other substance besides carbonic acid gas is formed by the burning. It only remains, therefore, to find out how much carbon is contained in 44 grains of carbonic acid gas: and it is found that 44 grains of this gas contain exactly 12 grains of pure carbon. From which we see that 12 grains of charcoal, 12 of black-lead, and 12 of diamond, each contain exactly 12 grains of pure carbon; in other words, that they are pure carbon.

VIII. THE AIR: NITROGEN.

If we burn a candle in a closed vessel, we shall find that after a few minutes it grows dim, and that after a few minutes more it goes out. The candle goes out when there is not sufficient oxygen left in the air (contained in the closed vessel) to enable it to burn. To find out about how much oxygen there is in the air is not difficult. Take a flat cork and float it in a basin of water; cover the upper surface with a layer of pounded chalk, and on this place a piece of phosphorus as large as a small pea. If this be touched with the heated blade of a knife, it will begin to burn. Directly it is alight, cover it with a glass bell-shaped vessel (a small glass ornament shade will do), letting the edges of the vessel dip well beneath the surface of the water. The vessel will be filled with white smoke from the burning phosphorus. This smoke is called **oxide of phosphorus.** Soon the phosphorus, just like the candle, will go out. In a few minutes the white phosphorus smoke will sink down and dissolve in the water; and it will then be found that about $\frac{1}{5}$th of the volume of the air in the vessel has disappeared, the water having risen in the vessel to take its place. This $\frac{1}{5}$th consisted of oxygen, which has disappeared because it united with the phosphorus to form the oxide of phosphorus smoke which dissolved in the water; the other $\frac{4}{5}$ths consist of a gas called nitrogen. Instead of using phosphorus we might use sulphur; in this case, instead of a white smoke being formed by the combination of the substance with the oxygen, a colourless gas called **oxide of sulphur** is produced. This gas dissolves in the water, and the water rises in the bell jar until ⅕th of the vessel is occupied by the liquid. Again, we might use spirits of wine or alcohol, as it is also called, instead of either phosphorus or sulphur. In this case we must float a little porcelain crucible or basin filled with the spirits of wine. The alcohol is, however, apt to go out before all the oxygen is burnt out, so that in this case the water will not rise quite so high as in the other cases.

About 100 years ago (in 1777), a great French chemist, Lavoisier, extracted the oxygen from the air in a somewhat different way. Instead of using phosphorus, sulphur, or alcohol, to unite with the oxygen, he used mercury (or quicksilver). He heated the mercury for several days in a closed vessel filled with air. The mercury took up the oxygen and was converted into a red powder. Nitrogen was left behind. Now the red powder, which consisted of mercury and oxygen, naturally weighed heavier than the mercury by itself. The extra weight was due to the oxygen. Thus Lavoisier was able to weigh the amount of oxygen in a given amount of air. And he found that if he took 100 grains of air, 23 of them were oxygen, and 77 were nitrogen. He also found that a candle would not burn at all in this nitrogen, and that if a mouse were put in a vessel full of this gas it would be suffocated. Hence he called the gas **azote** (Gr. *a*, not; *zoe*, life), the name by which it is still sometimes called in France. The name nitrogen signifies nitre-producer; for this gas is locked up in the nitre that is used in the manufacture of gunpowder. The gas is a little lighter than air; it has no taste, no colour, no smell, it has no action on a solution of litmus, it will not burn, nor will it allow other substances to burn in it, and in it an animal is unable to live. It does not poison the animal in any way; the creature dies of suffocation or want of oxygen. Thus we see that the properties of nitrogen are in the main negative. We shall presently learn, however, that when it combines with hydrogen, or with oxygen, or with hydrogen and oxygen, very important substances are formed. In the air the two gases, oxygen and nitrogen, are simply mixed together. And when we breathe we take both together into our lungs. There, some of the oxygen unites with carbon, and it thus comes out as carbonic acid gas. But the nitrogen is breathed out again unchanged. Thus we see that the air consists in the main of the two gases, nitrogen ($\frac{4}{5}$) and oxygen ($\frac{1}{5}$). Besides these gases there are smaller quantities of carbonic acid gas and vapour of water.

IX. THE CARBONIC ACID GAS AND WATER-VAPOUR IN THE AIR.

The chief constituents of the atmosphere are, as we have seen, nitrogen and oxygen. But other gases are also present in the air, though in smaller quantities. One of these is carbonic acid gas, of which there are from four to ten parts in every ten thousand parts of air. This is but a small proportion; but, as we have already seen, the carbonic acid gas is a most important constituent of our atmosphere, for it is the gaseous food of the vegetable world. To deprive plants of all carbonic acid gas would be to deprive them of one of the necessities of their life. The presence of this carbonic acid gas in the air may easily be shown by exposing to the air a vessel containing a little lime-water. The carbonic acid gas, acting on the surface of the lime-water, will form a scum or film of carbonate of lime.

Water-vapour is another important constituent of the atmosphere. Its amount varies very much with the temperature of the air. If the air be warm it can contain more moisture than it can contain when it is cold. This aqueous vapour is quite invisible. But when hot air, containing much of this invisible moisture, is cooled, the moisture becomes visible, in the form of the fine stream of which clouds and mist are made up. The air in this room is quite clear, but it contains a large quantity of vapour of water. If the air of the room, or any portion of it, be rapidly cooled, it cannot hold so much water-vapour in the invisible form. The invisible water-vapour will become visible water-drops. To show this we may take a test tube and pour in a little ether. Ether is a very volatile liquid—that is to say, it very readily passes into vapour. If now we blow a current of air through the ether, using a long india-rubber tube attached to a glass tube dipping into the ether, we shall cause some of it to evaporate very rapidly, and this will make the rest of the ether, and the tube which contains it, cold. After a few minutes, we shall see that minute drops of water dim the outside of the test tube. Those drops were originally present in the air; but they were then invisibly dissolved in the air. When, however, the warm air came in contact with the cold test tube it was cooled, and then could no longer hold so much invisible water-vapour. And so it was forced to deposit some of the moisture drop by drop upon the cold glass. Often during a cold night the blades of grass and cobwebs get so cold that the air is forced to deposit some of its moisture upon them, much in the same way as in the case of the cold test tube; and when we look out in the morning we see the fields glistening with these **dew-drops.** But as the air and the blades of grass get warmed by the sun, the dew-drops disappear, for the liquid water-drops assume again the condition of invisible water-vapour. If we watch a steam-engine, we can easily tell whether the air has much or little water-vapour contained in it. If it has much, it will not quickly dissolve the steam (from the engine), which will hang for a long while in the air. If it has little, it will quickly drink up the steam and hide it away as invisible water-vapour. When the air contains as much water-vapour as it can possibly hold, it is said to be **saturated.**

When air absorbs vapour of water it becomes lighter. This at first sight seems strange. One would have supposed that the air and water-vapour together would have weighed heavier than the air without the water-vapour. But this is not so. It has been ascertained by men of science that a given volume of gas (if the temperature and pressure do not change) always contains the same number of particles, no matter what the substance may be of which they are particles. Now water-vapour particles are much lighter than air particles. If we call the weight of the water-vapour particles 1, the average weight of the air particles is found to be 1.6, or rather more than 1½ times that of the water-vapour particles. Now suppose we have a space which can just contain ten particles. If the air be dry they are all air particles, and we have ten particles each weighing say 1½. The whole, therefore, weighs ten times 1½, that is to say 15. But if the air be moist, some of the particles will be water-vapour particles. Let us suppose that two of the ten particles are water-vapour particles. There are then eight particles weighing 1½, and two weighing only one. The whole will therefore weigh eight times 1½ or twelve, and twice one, or two, that is to say fourteen altogether. Dry air fifteen; moist air fourteen. The moist air is thus lighter than the dry air.

X. CLOUDS AND RAIN.

VAPOUR of water is invisible; but when vapour of water is cooled some of it is **condensed** into minute liquid drops, and these liquid drops are visible. If we watch a kettle, or a flask, in which water is being boiled, we shall notice that, where the vapour of water escapes from the spout, or from the mouth of the flask, it is invisible; but at a distance of an inch or so from the opening from which the vapour escapes, it is condensed into a little cloud of steam. So too on a rather larger scale, when we watch a steam-engine just starting from a station, at every puff a cloud of steam rises into the air; but this visible cloud does not make its appearance till the water-vapour has risen a foot or so above the funnel of the engine. Sometimes, especially when the air is damp, if we stand underneath the cloud of steam puffed forth from the engine, we shall find that liquid drops are falling like fine rain. The steam-engine, in fact, exemplifies, on a small scale, that which takes place on a large scale in nature. In warm air the vapour of water is quite invisible, but if the warm air be cooled the vapour of water is condensed into a cloud of steam particles. Ordinary clouds are simply clouds of very minute steam particles—very minute particles of liquid water. Each of these particles is a colourless and transparent droplet, but the multitude of particles which make up a cloud form a mass that is not transparent, but white and opaque. This is because the light cannot pass freely through a great number of separate transparent particles with air between them. It gets bent out of its course, and reflected hither and thither. For this reason, too, the foam on the crest of an ocean billow is white. We may exemplify this experimentally by taking a colourless crystal, or a piece of ice or glass. In the mass it is transparent. Now pound it up in a mortar. You get a white powder which is no longer transparent. It is indeed made up of transparent particles, but they form a white opaque mass, just as the minute transparent droplets make up a white opaque cloud. When clouds are forming very rapidly, and the air is damp, the minute water particles run together and collect into larger drops, and these drops as they fall pick up more cloud particles, and thus grow in size until they fall to the ground as rain-drops. Sometimes the raindrops freeze as they pass through a band of cold air. They are then converted into **hail-stones**. Sometimes the upper regions of the atmosphere are so cold that the clouds are formed, not of minute water-particles, but of minute ice-crystals. These, when they collect together and fall to the ground, form **snow-flakes**;

Here again the ice-crystals, of which the snow-flake is composed, are themselves transparent, but the entangled mass of these crystals, which constitutes a snow-flake, is not transparent but white and opaque.

The amount of rain which falls on the earth's surface at any particular place is called the **rainfall** of that place. It is measured in inches and portions of an inch. An "inch of rain" means this: that if we put out in the rain an open tank with vertical sides, the water will collect at the bottom to a depth of one inch. For measuring the rainfall, however, a special instrument is used, by means of which the rainfall may be measured to the fraction of one hundredth of an inch. The rainfall of any place varies according to the time of year and its geographical position. In the neighbourhood of London the rainfall in one year is about 20 inches: in the Khasia hills north-east of the delta of the Ganges, the rainfall in one year is more than 550 inches, most of which falls during four summer months. This rainfall has most important effects on the earth's surface. It enables the land to be clothed with vegetation: it fills our rivers and supplies our wells with water. And if we go out on a rainy day we shall see that it is washing some of the soil off the fields and roads into the ditches and streams and rivers. By the rivers it is carried down to the sea. So that the rain is constantly washing some of the surface of the dry land into the ocean.

XI. THE PRESSURE OF THE AIR.

A SHORT while before the death of Galileo, a great Florentine man of science, who lived during the first half of the seventeenth century, the engineers of Florence asked of him an explanation of a fact that puzzled them not a little. That fact was this. They found that, in the construction of some great fountains for the Grand Duke of Tuscany, some pumps, which were intended to raise water from a depth of between 40 and 50 feet, would not work. On looking into the question and making experiments with the pumps, Galileo found that the water in their tubes only rose to a height of about 32 feet. At that time the rise of water in a pump-barrel, when the piston is raised, was supposed to be due to the fact that Nature abhorred a vacuum. Galileo therefore said that Nature's abhorrence of a vacuum only extended to 32 feet. At the same time he requested a pupil, named Torricelli, to find out what he could about the matter. Torricelli was not content, as some men would have been, with sitting down quietly and thinking about it; he set to work and began to make experiments, and he soon found that 32 feet was the limit of Nature's abhorrence of a vacuum **for water only**. When he made an experiment with mercury, or quicksilver, a liquid more than thirteen and a half times as heavy as water, he found that Nature's abhorrence of a vacuum only extended to 30 **inches**. Let us try and imitate Torricelli's experiment. Let us take a tube about 33 inches long, closed at one end and open at the other, fill it with mercury, and close the open end with the thumb or finger. Inverting the tube, let us then dip the open end, still closed by the thumb, in a small basin of mercury, and finally remove the thumb, being careful that no air enters the tube. The mercury will sink in the tube, and will no longer fill it. There will be about three inches of open space at the top. This open space is the vacuum. We must be careful, however, to hold the tube vertically; for if the tube be gradually inclined the mercury will be seen gradually to occupy the whole of the tube. This shows that there is no air in the empty space to prevent the mercury rising. It shows, too, that the height, 30 inches, must be measured vertically. For it will be seen that the vertical height of the mercury remains the same, whatever the inclination of the tube. To make the same experiment with water in place of mercury, we should have to use a tube about 35 feet long; there would then be a vacuum of about two feet at the top of the tube. Torricelli had perhaps (for on this point we are not quite certain) been taught by Galileo that the air had weight, and therefore **exerted pressure**; it is quite probable that he knew that it was the pressure of the air that maintained the column of mercury in the one case, or of water in the other. The pressure of the air being limited in amount, the columns of mercury or water which it could support would be limited in length.

About the same time (1644) Pascal heard of the experiments of Torricelli, and repeated them with perfect success. Soon afterwards he heard, also, of Torricelli's explanation (an explanation which he had himself independently proposed), and he determined to test or verify it by experiment. He saw that if the column of mercury was kept up by the pressure of the air, it would vary in length at different heights above the sea level. He therefore begged his brother-in-law (for he was too ill to carry out this experiment for himself), to take a closed glass tube, four feet long, to the top of the Puy de Dome, a mountain in Auvergne, nearly 6000 feet high. His brother-in-law did so. He first ascertained by experiment how long a column of mercury the pressure of the air maintained at the foot of the mountain; and then when he arrived at the top of the mountain he ascertained by experiment how long the column was there. He found that the column was about six inches shorter on the mountain top than at the mountain foot. This is exactly what Pascal had said would happen; and it showed that the explanation which had been given of Nature's abhorrence of a vacuum was correct. The facts were due to the pressure of the atmosphere.

The instrument used for measuring the atmospheric pressure is called a **Barometer.** It is a tube, like that used in Torricelli's experiment, in which a column of quite pure mercury is supported by the pressure of the atmosphere. The average level of the mercury in the tube at the sea-level, is 760 mm., or 29·92 inches. That is to say, the average atmospheric pressure at the sea-level will support a column of mercury 760 mm., or 29·92 inches long. The pressure of the atmosphere is, however, subject to changes. Sometimes the pressure becomes less. It will not then support so long a column of mercury. The "barometer falls"; that is to say, the level of the mercury in the tube sinks below 760 mm. Sometimes the pressure increases. The mercury then rises, and stands at a height of more than 760 mm. in the tube. These changes of pressure often precede changes of weather. Hence the barometer can be used as a "weather glass." From all this we see that the air or atmosphere which rests on the surface of the earth has weight. The weight of the great mass of air over our heads is very considerable, and exerts a pressure of about **fifteen pounds** on every square inch, or 1·03 kilometres on every square centimetre. But this pressure is exerted in all directions. Suppose we cut a square piece of paper, one square inch in area, and running a pin right through the middle fix the pin in the table. There is a pressure of fifteen pounds on the top of the square of paper; but the paper is not pressed down on the table, for there is a similar pressure of fifteen pounds upwards on the under side of the piece of paper. The pressure is in fact transmitted in all directions; and this is why we do not feel the atmospheric pressure as a great weight on our heads and shoulders. Now if we take a barometer to the top of a mountain, thus performing Pascal's experiment, we shall not have so great a mass of air above us as we have at the sea level; the pressure will therefore be less, and the mercury column in the barometer tube will be below 760 mm. It is thus possible to tell the height of a mountain by means of a barometer. For if we know how much less the pressure of the air is on the mountain top than it is at the sea level, we can calculate how high the mountain top is above the sea level. If instead of taking the barometer tube to the top of a mountain we take it to the bottom of a deep mine, we shall find that the mercury stands at a level above 760 mm. This is of course because there is a longer column of air pressing downwards, and this longer column with its greater pressure supports a longer column of mercury.

XII. THE THERMOMETER.

If we take a test tube, and fill it quite full of water, and then fix into the mouth of the test tube a cork, through which a thin straight glass tube passes—open at both ends—some of the water will be forced a little way up this thin tube. The test tube and the lower part of the thin glass tube will be full of water. If now we heat, by means of a spirit lamp, the water in the test tube, we shall see that the water rises in the thin glass tube. Why does it rise? Because the water **expands** when it is heated. Now let us stop heating the water. The water cools; and as it does so its level gradually sinks in the fine glass tube. It sinks because, as the water cools, it **contracts.** When it is cold again we may make a mark on the thin glass tube at the height at which the water stands. We may then heat it for a little while: and put another mark: heat again and put a third: and so on as the temperature rises. These marks will show how high the water stands at certain temperatures. Each mark will stand for a certain temperature of the water in the tube. We have thus constructed a very rough instrument for measuring temperature—a very rough **thermometer.** In thermometers in common use the liquid is not water but mercury, or sometimes spirits of wine coloured red. Such a thermometer consists of a tube of very fine bore; that is, a tube in which the hole or bore is not much wider than a thick hair.

At the end of the tube there is a bulb. The tube is filled with mercury, an operation of some little difficulty which cannot be described here, and then the bulb is placed in hot oil. This makes the mercury in the bulb expand. The mercury now fills the whole of the bulb, and the whole of the fine bore of the tube. The open end is then closed by melting the glass, and the bulb is taken out of the oil. As the mercury cools it sinks in the tube, leaving an empty space, or vacuum, above it. The bulb of the tube is now placed in ice, and a mark is made on the tube to mark the level of the mercury. The bulb is then taken out of the ice and placed in the steam that rises from boiling water. Another mark is made on the tube to show the level of the mercury at this temperature. The space between the two marks is now divided into 100 equal divisions called **degrees.** The mark which shows the level of the mercury in ice is marked 0°, (the small sign to the right of the figure standing for degree). The mark which shows the level of the mercury in steam is numbered 100°. The divisions between the two are numbered from 1° to 100°. Similar divisions are marked off below 0° and above 100°. Those below 0° are numbered downwards and marked –1, –2, –3, and so on.

The thermometer just described is known as the Centigrade thermometer, and is so called because the space between the freezing point of water and the boiling point is divided into 100°. On such a thermometer the freezing point is zero. The Centigrade thermometer is nearly always used by scientific men in England, and is in general use on the continent of Europe. The thermometer in general everyday use in England is called the Fahrenheit thermometer. In it the freezing point of water is not marked zero, but 32°. The zero on this scale corresponds to the temperature obtained by mixing salt and snow. As this temperature was once supposed to be the lowest obtainable, it was made the zero or starting point so as to avoid minus readings. Temperatures far below this, however, have now been reached. The boiling point of water on the Fahrenheit scale is 212°, so that the space between the freezing point and boiling point in this thermometer is divided into 180 degrees. One hundred and eighty Fahrenheit degrees are therefore equivalent to one hundred Centigrade degrees.

It is useful to have a rule for converting Centigrade readings into Fahrenheit, and Fahrenheit into Centigrade.

To convert C.° into Fah.° Multiply the number of degrees C. by 9, divide the product by 5, and add 32; or

$$F. = \tfrac{9}{5} C. + 32.$$

To convert Fah.° into C.° Subtract 32 from the number of degrees Fah., and then multiply by 5, and divide the product by 9; or

$$C. = \tfrac{5}{9} (F. - 32.)$$

The reason of these rules will best be seen by making a diagram of the two thermometers side by side. It will then at once be seen that 9° Fah. are equal to 5° C., and that a Centigrade degree is therefore $\frac{5}{9}$ of a Fahrenheit degree. The addition or subtraction of 32 will at first appear confusing. But a little consideration will make it clear that, in the conversion of Centigrade into Fahrenheit, multiplying by 9, and dividing by 5, gives the number of degrees—

not above zero, but above the freezing point. But on the Fahrenheit scale the freezing point is already marked 32. The number we have obtained is, in fact, the number of degrees above 32°. This 32 must therefore be added to obtain the actual reading on a Fahrenheit thermometer. So, too, in the conversion of Fahrenheit into Centigrade, we can only make use of the fact that 1° C. is equal to $\frac{9}{5}$ Fah., on condition that the Fahrenheit and Centigrade degrees start fair from the same point. But the Fahrenheit reading starts at a point 32 degrees below the starting point of the Centigrade. These 32 degrees must therefore be subtracted from the Fahrenheit reading, and then a fair start may be made from the freezing point of water in each case, and we may proceed to multiply by 5 and divide by 9.

NOTE.—A number of examples should be worked by those who have sufficient knowledge of arithmetic, and the results obtained compared with the actual readings of Centigrade and Fahrenheit thermometers.

XIII. CHANGES OF VOLUME IN THE AIR.

WE know from everyday experience that the temperature of the air is subject to changes; at night, for example, it is cooler than in the daytime. We have learnt, too, that the atmospheric pressure is subject to variation. Let us in this lesson see if we can learn anything by experiment concerning the behaviour of air under the influence of these changes of temperature and pressure. Let us take a glass flask fitted with a cork, through which a small glass tube passes, and holding it in such a manner that the end of the glass tube dips under the surface of some water in a tumbler, let us heat the flask by means of a spirit lamp. We shall find that bubbles of air rise rapidly from the end of the glass tube to the surface of the water. What does this show? The flask to begin with is full of air. We heat the flask, and some of the air escapes. This shows that the heat we apply to the flask expands the air within the flask. When we cease heating the flask the bubbles cease rising from the end of the tube. And as the flask cools water is gradually sucked up the tube. What does this show? That as the heated air in the flask cools, its volume becomes less. The air contracts on cooling. Thus we learn by experiment that the air expands and contracts under the influence of changes of temperature. And not only air, but every gas **expands when it is heated, and contracts when it is cooled.**

We have now to see what effects follow on change of pressure. Let us take some air in a glass cylinder closed with a tightly-fitting piston. If we press down the piston we **diminish** the volume of the gas, squeezing it into a smaller space. If, on the other hand, we lift the piston, we cause the gas to **expand** or increase in volume. Thus we see that increase of pressure diminishes the volume of the gas, while decrease of pressure allows of the expansion of the gas. If we cannot obtain a glass cylinder fitted with a piston, there is another way in which the fact, that increased pressure produces diminished volume, may be shown. Let us take a tube shaped like a U, one end of which is open, the other being closed. Let us pour in a little mercury. It will occupy the bottom of the U tube, and a certain volume of gas will be enclosed between the mercury and the closed end of the tube. Mark the height at which the mercury stands in that tube. Now pour in at the open end of the tube more mercury. The column of mercury that collects on this side of the U tube exerts a pressure on the volume of air enclosed in the other side of the tube, and as this column (and therefore the pressure) increases the volume of enclosed gas diminishes. Now, instead of using an U shaped tube, let us take a long straight tube like that used for the Torricellian experiment, and holding it open end upwards, fill it with mercury within four or five inches of the mouth of the tube. Let us then close the open end with the thumb or forefinger, and inverting the tube, hold the open end under the surface of some mercury in a basin. On carefully removing the thumb or finger the mercury will be seen to sink in the tube, the air above the mercury expanding. Why does the air expand? Because the weight of the mercury in the tube takes off some of the atmospheric pressure which was before compressing the gas. Let us compare this experiment with that with the U shaped tube. Remember that the particles of a gas are always trying to fly apart from each other. The air enclosed in the U tube is, to begin with, at the atmospheric pressure. More mercury being then poured in gives an increased pressure, and squeezes the air particles closer together. The air enclosed in the straight tube is also, to begin with, at the atmospheric pressure. But when the tube is inverted and the thumb removed, the mercury, instead of squeezing the air particles together, tends to pull them apart, or rather allows them to fly further apart from each other, and thus allows the air in the tube to expand.

XIV. OXYGEN.

WHEN mercury or quicksilver is heated for a long time in the air (see lesson viii.) it combines or unites with some of the oxygen in the air, and is converted into a red powder. This powder is called **oxide of mercury. In this oxide, as in every true chemical compound, the substances are united in proportions which are always the same.** Two hundred parts of mercury always unite with sixteen parts of oxygen, neither with more nor less. Now if we take this mercury oxide and heat it in a stout glass tube—making it as hot as we can—we shall split it up or decompose it into mercury and oxygen. The mercury will collect as minute drops in a ring round the inside of the cooler part of the glass tube. The oxygen will pass as a gas out of the mouth of the tube. We can find out or test for the presence of the gas, by putting a match which has been blown out—but the end of which is still aglow—into the open end of the tube. It will burst into flame with a slightly explosive noise. We can make oxygen in somewhat greater quantity, and more easily, by heating a substance called **potassium chlorate**, with which a little black oxide of a metal called manganese has been mixed. We may do this in a glass vessel fitted with a cork, through which a glass tube passes. We may connect this tube, by means of india-rubber tubing, with another glass tube, dipping under the surface of some water in a basin or trough; and we may then fill some tall jars with water and hold them in succession, mouth downwards, and still full of water, over the end of the glass tube. On heating the potassium chlorate oxygen will be set free from this substance, in which it is, as it were, locked up. The gas will pass down the tube and bubble up into the glass jars, pushing out the water until the jars are full of the gas. In this way we may collect four jars-full for experiment.

Leaving a little water in the bottom of each, we may in No. 1 burn a taper. It burns with a brighter light than usual. In No. 2 we may burn a little sulphur. It burns with a beautiful violet light. In No. 3 we may burn some phosphorus. The light is so bright as to be dazzling. In No. 4 we may burn some iron wire (having first dipped the end in burning sulphur). It burns with a number of bright sparks. If after burning the first three substances for a while we shake up the water in the jars, and then add some blue litmus, the blue will turn to red. This shows that **acids** have been formed. Hence the name oxygen or acid-producer (Gr. *oxus*, acid; *gennao*, I produce). The iron, however, will have become a **brittle** globule of **oxide of iron**. Of these four substances the taper, the sulphur, and the phosphorus, will burn in the air as well as in oxygen; but they do not then burn so long or so brightly. Iron wire will not burn in the air under ordinary circumstances; but if we drop some iron filings into the flame of a spirit lamp, many of them will catch fire and burn with bright flashes. In all these cases in which substances burn in the air, it is the oxygen with which they chemically unite. The nitrogen takes no part in the burning, and only serves to dilute the oxygen, just as spirits of wine may be diluted with water. There are, however, some substances which unite with the oxygen of the air slowly and quietly, without giving rise to the bright light we are accustomed to in ordinary burning. Phosphorus is one of these. If we hang a piece of phosphorus in the air it slowly wastes away, a white smoke (consisting of oxide of phosphorus) being formed. We cannot, therefore, keep phosphorus in the air; it must be kept under water, which excludes the air. The metal sodium is another substance which unites with the oxygen of the air at the ordinary temperature, quietly and without burning. If we cut a piece of sodium a bright metallic surface is exposed; but this at once tarnishes with a film of the oxide of the metal. Sodium has to be kept in naphtha or wood spirit. It cannot be kept like phosphorus under water. We shall presently know why.

To sum up concerning oxygen: this gas is colourless and invisible; it supports combustion, but will not burn itself; it is without smell when quite pure; and is slightly heavier than the air of which it is an important constituent, one fifth of the air by volume being composed of this gas.

XV. ELEMENTS AND COMPOUNDS.

It is part of the business of science to find out of what materials all substances are composed. When a substance is composed of two or more different kinds of substances it is called a **compound**: but when a substance contains only one kind of substance it is called an **element**. In language most written **words** are like compounds, because they contain two or more different kinds of letters: but the words **I** and **a** are like elements, for they each contain only one kind of letter. Carbonic acid is a compound, because it contains two kinds of substance, oxygen and carbon. Oxide of mercury and carbonate of lime are also compounds. But oxygen, carbon, nitrogen, phosphorus, are elements because they each contain only one kind of substance. There are in the world more compounds than can be numbered—nearly everything we see around us being a compound substance. But all these are built up out of between sixty and seventy elements; and nearly all out of about ten or twelve elements. Some of the elements most commonly known are the metals, such as gold, silver, copper, lead, tin, and iron. At the end of this lesson a list of some of the commoner elements is given.

It is perfectly impossible to convert one element into another element—to convert, for instance, lead into gold. This is what the alchemists of old tried to do and always failed. We must not suppose, however, that, because they tried to do that which we now know to be impossible, they were foolish so to try. On the contrary, it was through their early attempts that people began to learn anything of these matters. They did not know that gold and lead were elements, and that therefore it was impossible to convert one into the other. They seemed, in fact, to be able to convert iron into copper. In an old writer, for example, we read to this effect: "From iron copper can be got by natural means, for instance, by an acrid ley (liquid) from Hungary, which gives to it such a metallic colour that it is converted into the best copper." The acrid ley was a liquid containing copper, and what takes place when iron is put into the liquid is, that the iron dissolves and copper takes its place. But if we dip a knife into the "acrid ley" it comes out so covered with a thin film of copper that it seems to have been actually converted into copper. There seems to be little doubt that the great Sir Isaac Newton laboured at alchemical experiments, and tried to discover the secret by which gross metals might be rendered noble. These alchemists of old thought—or some of them did—that the metals were compounds of what they termed sulphur and mercury. Gold and silver, they believed, contained a very pure mercury. Lead contained more sulphur. They imagined, therefore, that if they could only get rid of the excess of sulphur in the lead, this metal would be converted into gold. And so they tried in a great number of ways to get rid of the sulphur. They did not succeed because lead is an element; and it is therefore impossible to obtain from lead any other metal than lead. But by all their trials they learnt a great number of things which were very well worth learning, and collected a great number of facts which have formed the foundation on which the modern science of chemistry is based.

SOME ELEMENTS WITH THEIR SYMBOLS AND ATOMIC WEIGHTS.

ELEMENT.	SYMBOL.	AT. WT.
Bromine	Br	80
Calcium	Ca	40
Carbon	C	12
Chlorine	Cl	35·5
Copper (Cuprum)	Cu	63
Fluorine	Fl	19
Hydrogen	H	1
Iodine	I	126·5
Iron (Ferrum)	Fe	56
Lead (Plumbum)	Pb	206·5
Magnesium	Mg	24
Manganese	Mn	55
Mercury (Hydrargyrum)	Hg	200
Nitrogen	N	14
Oxygen	O	16
Phosphorus	P	31
Potassium (Kalium)	K	39
Sodium (Natrum)	Na	23
Silicon	Si	28
Sulphur	S	32
Zinc	Zn	65

XVI. COMPOUNDS AND MIXTURES.

THERE is a great difference between simply mixing things together and causing them so to unite as to form a chemical compound. Take some iron filings and sulphur. You may mix them together in any proportion: the mixture will be a mere mixture and nothing more; at any moment you may, by means of a magnet which attracts the iron, pull out all the metal and leave the sulphur. Or you may, by means of a liquid called carbon disulphide, dissolve out all the sulphur and leave the iron behind. But now suppose you heat a mixture of sulphur and iron in certain proportions in a glass tube: you will see it soon glow a deep red. Then if you examine the stuff with a strong glass, you will no longer be able to detect, as you could detect before, the different particles lying side by side. All is now of a uniform grey colour. A magnet will no longer enable you to separate the iron from the sulphur; nor will carbon disulphide dissolve out the sulphur and leave the iron behind. A chemical compound of the iron and sulphur has been formed: and a chemical compound always has **new properties**: properties that differ from those of the elements of which it is composed. We have already seen some examples of this: we have seen that carbonic acid gas differs from both the carbon and the oxygen of which it is composed: that in the white material of which the candle is composed, the carbon and hydrogen have assumed a form very different from that of the black solid carbon or the aseous hydrogen; and, once more, that the liquid water drops differ from the gaseous oxygen and hydrogen of which they are composed. Let us, however, take one more example. Here is a clean bright strip of copper. The metal is reddish in colour and tough: you cannot reduce it to powder by hammering it. Now put it and some sulphur at the bottom of a glass test tube, and heat. Soon, as in the case of the iron, the copper glows red. Examine it when it is cool, and you will find that it is no longer like copper; it is not reddish, but steel grey in colour and sparkling. And if you hammer it, you at once reduce it to powder. It is a chemical compound. The most important thing to notice about a chemical compound, however, is this: that the union takes place in **definite proportions by weight.** For however the properties of bodies may change during chemical union or chemical decomposition, **one thing remains constant, and that is the weight.** Fifty-six parts by weight of iron unite with thirty-two of sulphur; and sixty-three parts by weight of copper unite with thirty-two of sulphur. Carbonic acid gas is also a chemical compound, and the proportions of carbon and oxygen are therefore definite and invariable. Twelve parts by weight of carbon unite with thirty-two parts by weight of oxygen, no more and no less. Carbonate of lime, too, is a chemical compound, and has a perfectly definite composition, there being forty-four parts of carbonic acid gas to fifty-six parts of quicklime. We cannot alter the proportions of the elements in any of these compounds. But the air is not a chemical compound. The oxygen and nitrogen are simply mixed together. This may be seen by the following facts.

1. When a chemical compound is produced by the union of elements, its formation is accompanied by a change of temperature: we see this in the burning candle and the glowing copper. But when four parts of nitrogen are mixed with one part of oxygen, there is no change of temperature. The mixture is, nevertheless, in all respects like air.

2. The properties of a chemical compound are altogether different from the properties of its constituents. The properties of air are, however, a mean between those of its constituent gases.

3. The proportions of the elements in a chemical compound are absolutely constant. The proportions of oxygen and nitrogen in the air, though very nearly invariable, are not absolutely constant.

4. By dissolving a portion of a chemical compound in water, no alteration in the proportions of the elements is produced. By shaking up air with water, the proportions of the oxygen and nitrogen which dissolve, are nearly one of oxygen to **two** of nitrogen, instead of one of oxygen to **four** of nitrogen. That is to say, solution produces an alteration of the proportions of the elements.

5. For those who have read Lesson XVIII. we may add a fifth point of difference. In a chemical compound the proportions of the elements are simple multiples of certain numbers called the atomic weights of the elements. In air the oxygen and nitrogen are not in simple multiples of their atomic weights.

XVII. HYDROGEN.

A CHEMICAL compound has very different properties to those possessed by the elements of which it is compounded. Water is a compound: and no amount of study of the properties of this compound would tell us anything of the elements by the union of which it is formed. For a long time, indeed, water was regarded as an element, that is to say, a substance which could not be decomposed into two or more simpler substances. And with the discovery of the true composition of water the names of Henry Cavendish and James Watt (1781) are associated. They found out, independently, that water is composed of the two gases which we call oxygen and hydrogen. The way in which the exact composition of water is ascertained will be described in Lesson xx. Unlike nitrogen and oxygen, hydrogen is not found in the free and gaseous state in the atmosphere. Great quantities of the gas seem, however, to be sometimes emitted from the mouths of volcanoes. The hydrogen thus emitted at once unites with the oxygen of the air, the water produced by the combustion forming clouds of steam. There are several ways of preparing hydrogen. Perhaps the best is to put some pieces of the metal zinc into a glass vessel, like that used for making carbonic acid gas, and pour upon these water and sulphuric acid. Hydrogen will be given off rapidly, and may be led off by means of a tube. The gas may be collected in test tubes or glass cylinders over water, for water does not dissolve hydrogen. If we hold the mouth of a test tube full of hydrogen to the flame of

XVIII. ATOMS AND MOLECULES.

A CHEMICAL compound consists of two or more elements united in definite proportions by weight. Carbonic acid gas consists of carbon and oxygen in the proportions of twelve of carbon to thirty-two of oxygen. Water consists of sixteen of oxygen to two of hydrogen; oxide of mercury of two hundred of mercury to sixteen of oxygen. Carbonate of lime consists of three elements; for the quicklime that is left when the carbonic acid gas has been driven off by heat, is an oxide of a metal called calcium. In carbonate of lime, then, the proportions are forty of calcium, twelve of carbon, and forty-eight of oxygen. The white smoke formed when phosphorus is burnt in the air consists of phosphorus and oxygen, in the proportions of sixty-two of phosphorus to eighty of oxygen. All these facts have been found out by the careful experiments of men of science. If now we look at the quantities of oxygen we find that, in the water and the oxide of mercury there are sixteen parts; in the carbonic acid gas there are twice sixteen, or thirty-two parts; in the carbonate of lime, there are three times sixteen, or forty-eight parts; and in the oxide of phosphorus there are five times sixteen, or eighty parts. Sixteen or some multiple of this number occurs in each case.

The smallest visible quantity of any of these compounds contains a great number of minute particles which are called molecules. And each molecule of a compound must consist of still minuter particles of the elements of which the compound is composed.

A molecule of carbonic acid gas for example must consist of minute particles of carbon and oxygen; a molecule of water, of minute particles of oxygen and hydrogen. These minute particles of elements are called **atoms**. To be exact we may here define an atom as the smallest quantity of an element which can enter into or be expelled from a chemical compound. A molecule of any substance is the smallest group of atoms of that subject which is capable of separate existence. A molecule can only be broken up into its constituent atoms by chemical means. Hydrogen is the lightest body known, and the atom of hydrogen is the lightest atom with which the chemist is acquainted. Men of science, therefore, always speak of atoms of the other elements as so many times as heavy as the hydrogen atom; the oxygen atom is sixteen times as heavy as the hydrogen atom; that is why the number sixteen, or a multiple of sixteen, occurs in all the cases given above; each sixteen stands for one atom, for the numbers there given represent the proportional weights of the elements in one molecule of the compound. And we must remember that the proportions of the elements present in one molecule are the same as the proportions of the elements in a collection of molecules, since all the molecules are alike. The carbon atom is twelve times as heavy as the hydrogen atom; the mercury atom two hundred times; the calcium atom forty times; the phosphorus atom thirty-one times, and so on. Looking back now to the proportions of the elements in those compounds which are given at the beginning of this lesson, and making use of the weights of the atoms just given, we find that the carbonic acid molecule contains one atom of carbon, and two atoms of oxygen; that the water molecule contains one atom of oxygen and two atoms of hydrogen; the molecule of carbonate of lime one atom of carbon, one of calcium, and three of oxygen; the molecule of oxide of mercury one atom of mercury and one atom of oxygen; and the molecule of oxide of phosphorus two atoms of phosphorus and five atoms of oxygen.

These results will be seen more clearly if we tabulate—

Carbonic acid gas. Carbon, 12 (= 1 atom); Oxygen, 32 (= 2^{ce} 16 or 2 atoms).

Water. Oxygen, 16 (= 1 atom); Hydrogen, 2 (= 2^{ce} 1 or two atoms).

Carbonate of lime. Carbon, 12 (= 1 atom); Calcium, 40 (= 1 atom); Oxygen, 48 (= 3 × 16 or 3 atoms).

Mercury oxide. Mercury, 200 (= 1 atom); Oxygen, 16 (= 1 atom).

Oxide of phosphorus. Phosphorus, 62 (= 2^{ce} 31 or 2 atoms); Oxygen, 80 (= 5 × 16 or 5 atoms).

XIX. THE HYDROGEN IN WATER.

When sulphuric acid is poured upon metallic zinc hydrogen is formed. The question now arises, where does the hydrogen come from. Zinc is an element; it is certain, therefore, that the hydrogen does not come from the zinc. It must come, then, from the sulphuric acid. The accurate experiments of men of science have, in fact, proved that sulphuric acid is a compound, containing two parts by weight of hydrogen, combined with thirty-two parts by weight of sulphur, and sixty-four parts by weight of oxygen. Now, we have seen that the weight of the hydrogen atom is one, and that of the oxygen atom is sixteen. The weight of the sulphur atom is thirty-two. From this we see that there are in sulphuric acid two atoms of hydrogen, one atom of sulphur, and four atoms of oxygen. What takes place when the sulphuric acid is poured on the zinc is this. One atom of zinc throws out and takes the place of two atoms of hydrogen in the sulphuric acid, joining itself to the atom of sulphur and the four atoms of oxygen to form a new compound called zinc sulphate. If we put some concentrated sulphuric acid at the bottom of a test tube, and then drop in one or two pieces of zinc, a few bubbles of hydrogen will be given off at first, and then the action will cease. It ceases because a film or layer of zinc sulphate forms over the metallic zinc, and as this does not dissolve in the concentrated sulphuric acid, it prevents the acid from getting at the rest of the zinc. Now very carefully add some water. The hydrogen is given off rapidly until the zinc has all dissolved. The water dissolves the film of zinc sulphate as fast as it forms, and thus enables the sulphuric acid to get at fresh portions of the zinc. Notice that the chemical action makes the liquid in the bottom of the test tube quite hot. Put the test tube on one side to cool, and when it is quite cold you will see that, if too much water has not been added, at the bottom there is a white crystalline solid. That solid is the zinc sulphate.

Water also contains hydrogen. A molecule of water contains one atom of oxygen combined with two atoms of hydrogen. If, however, we put a piece of zinc in water it will not throw out and take the place of the hydrogen in the water, as it does in the sulphuric acid. But if we throw upon water a small pellet, as large as a pea, of the metal sodium or the metal potassium, both of which are lighter than water, and therefore float, the metal will throw out and take the place of the hydrogen in the water. Let us try the sodium first. Directly the pellet of sodium falls upon the water, it assumes the form of a bright metallic ball, which rolls rapidly about on the surface, buoyed up by the hydrogen which is given off from its under side. As it rolls about it gets rapidly smaller and smaller, until at last it darts about as a minute speck, and then disappears, perhaps with a little flash of yellow light. Now let us try the potassium. Directly the pellet of potassium falls upon the water it bursts into flame, and the globule of burning metal rolls about on the water for a very short time, and then goes out. The flame is violet and gives rise to a cloud of smoke. After the metal has ceased to burn, there still rolls about on the surface of the water a colourless, transparent globule, which after a moment or two suddenly explodes with a crackling noise. What takes place in these cases is this: when the metal falls on to the surface of the water it at once takes the place of and expels some of the hydrogen of the water. The chemical action is in each case accompanied by heat; but the heat is greater in the case of potassium, so that the hydrogen given off catches fire, and this causes the globule of metal to catch fire also, and to burn away rapidly. The white smoke which rises from the burning metal is some of the oxide of the metal produced by its combustion.

If we throw sodium on to **hot** water, or on to wet blotting paper, it too catches fire and behaves just like the potassium. But the colour of the flame is not violet, but yellow.

The molecule of water contains one atom of oxygen combined with two atoms of hydrogen. When sodium or potassium acts on the water, one atom of the metal throws out or displaces one atom of hydrogen in the water-molecule, uniting with the other atom of hydrogen and the atom of oxygen to form a new compound, called a **hydrate** of the metal, the name hydrate being derived from the Greek word for water (*hudor*). The colourless globule that floats for a moment on the water after the potassium has ceased to burn, is the hydrate of potassium. The hydrate dissolves in the water. To show its presence we may add a little **litmus** solution which has been turned red by an acid. The reddened litmus becomes blue again. Those substances which turn reddened litmus blue again are called **alkalies.** The hydrates of potassium and sodium are therefore alkalies.

If a pellet of sodium be wrapped up in a piece of fine wire-gauze, and the whole be thrown into water, the heavy wire-gauze will make the sodium sink to the bottom. There it will decompose the water, and the hydrogen will be able to escape to the surface through the meshes of the gauze. The rising bubbles of the gas may be caught in an inverted test tube full of water. The gas will be found to have all the properties of hydrogen.

Several metals which will not decompose water when they are cold, will decompose steam when they are red-hot. Iron is one of these; and if a current of steam be passed over red-hot iron in a gun-barrel, or some such tube, the iron will throw out and displace all the hydrogen in the water molecules, setting free the hydrogen, and uniting with the oxygen to form a new compound called iron oxide.

XX. THE COMPOSITION OF WATER.

WHEN hydrogen burns in the air water is formed. And wherever we go on the surface of the earth water is found in greater or less abundance. It falls from the clouds as rain; it rises from the earth in springs; it exists in the great oceans. But in the ocean, in springs, and even in rain, it is not quite pure water. This is because water has the power of dissolving many things, such as oxygen in small quantities, carbonic acid gas, and, by means of the carbonic acid, limestone. Nitre, washing soda, and common salt, are other things which dissolve in water. These things make the water impure. We can, however, by proper means, such as distilling the water (see Lesson XXXI.) get rid of these impurities; and then we find that quite pure water always consists of the **same** two elements, oxygen and hydrogen, united in the **same** proportions. If we pass a current of electricity through water the water is split up into its elements. It is well to add a little sulphuric acid to the water to enable the electricity to pass more easily. The acidulated water may be placed in a funnel-shaped vessel in the bottom of which is a cork. Through the cork pass two platinum wires, which are attached to two flat pieces of platinum foil, immersed in the acidulated liquid. If now we join the other ends of the platinum wires, outside the vessel, to the copper wires coming from a small battery, we shall see that bubbles of gas rise from each of the two pieces of platinum foil. The bubbles may be collected in inverted test tubes full of water. We shall soon see that twice as much gas rises from one of the pieces of platinum foil as from the other. On testing these gases we shall find that the gas in one test tube (that in which there is a double amount of gas) burns with a blue flame. It is, in fact, hydrogen. The gas in the other test tube rekindles a glowing match. It is, in fact, oxygen. In an actual experiment we may at first find that there is rather more than twice as much hydrogen. This is because some of the oxygen dissolves in the water. Two volumes of hydrogen to one volume of oxygen is therefore the composition of water by volume.

Now, volume for volume, oxygen is sixteen times as heavy as hydrogen. We may therefore conclude that the composition of water by weight is sixteen parts of oxygen (since that is the relative weight of one volume), and two parts by weight of hydrogen, each volume weighing one. Let us now see how this may be verified. It is found that if hydrogen gas be passed over heated copper oxide—a compound of one atom of copper and one atom of oxygen—this copper oxide gives up some of its oxygen to the hydrogen. The hydrogen uniting with this oxygen, which it steals from the copper oxide, forms water. This water may be caught in a tube, which contains a chemical substance (calcium chloride) which readily absorbs moisture. Suppose, then, we place some copper oxide in a tube, and connect this tube with a second, in which is some calcium chloride to catch the water; and then making the copper oxide red hot, pass through the tubes a stream of hydrogen gas, which has been dried by passing it through a tube containing calcium chloride. The copper oxide will give off some of its oxygen, thus becoming lighter, and the calcium chloride will absorb all the water formed, thus becoming heavier. Let us notice in passing, that the copper oxide which before the experiment was black, becomes red. This is because some of the copper oxide has lost its oxygen, and become metallic copper. Now the question is, **how much** lighter the copper oxide becomes, and **how much** heavier the calcium chloride becomes. To find out this we must carefully weigh the copper oxide, in its tube, before the experiment and after the experiment; and the calcium chloride, in its tube, before the experiment and after the experiment. Here are some weighings—

I.

Copper oxide at first	228
„ „ after experiment	204
Difference due to loss of oxygen ..	24

II.

Calcium chloride after experiment	777
„ „ at first	750
Difference due to gain of water	27

III.

Water formed	27
Oxygen in that water	24
Difference due to the hydrogen in the water	3

From these weighings we see that in twenty-seven grains of water there are twenty-four grains of oxygen, and three grains of hydrogen. All pure water has exactly this composition by weight, namely, eight parts by weight of oxygen for every one part of hydrogen.

To sum up. We see, therefore, that twice as much hydrogen as oxygen, **by volume**, unite to form water; but that eight times as much oxygen as hydrogen, **by weight**, unite to form water. It is just as if we had two cubic inches of feathers, and one cubic inch of lead. We should then have a greater bulk of feathers, but a greater weight of lead. So here we have a greater volume of hydrogen, but a greater weight of oxygen.

XXI. ICE.

If, on a cold winter's evening, we fill a common medicine bottle quite full of water and cork it up tightly, tying down the cork with strong string, and then leave the bottle all night out of doors in the frosty air, we shall find next morning, if the cold has been severe enough to freeze the water, that the bottle has burst. Another way of making the same experiment is to place the bottle in a mixture of pounded ice (or snow) and salt, or in the freezing mixture described at the end of this lesson. Why does the bottle burst? Because the water when it is converted into ice expands, or increases in volume, and the bottle is not large enough to hold the mass of ice formed, nor strong enough to resist the expansive force. We know well, and if we do not we may ascertain by experiment, that ice floats on the surface of water. If water did not expand, but contracted on freezing, ice would not float, but would sink to the bottom. Since it expands it occupies more space than water; and a cubic inch of ice weighs less than a cubic inch of water; and the ice weighing less than the water floats on its surface. When a piece of ice floats in a tumbler of fresh water, about one twelfth of its mass is above the surface, eleven twelfths being immersed in the water. If we place a piece of ice in a tumbler of warm water, and hold it up to the light, we shall see a stream of cold water sinking from beneath the ice. This very simple experiment shows that the ice is melting in the warm water, and that the cold water formed from the melting ice is heavier than the warm water, through which it sinks towards the bottom.

The ice-bergs, about which we read in accounts of Arctic voyages, are masses of ice broken off from the glaciers, which in Arctic countries reach the sea. Glaciers are huge masses of ice which fill the valleys in those parts of the earth where it is so cold that the snow which falls during the winter does not melt away and disappear during the summer. The glacier is made of snow which has been squeezed and soldified into ice. Its surface is snow-white and crunches under one's feet like sugar. But here and there, where the ice is cracked, and we can see into the heart of the glacier, the ice appears blue or bluish green. These glaciers are always moving slowly downwards and seawards. In 24 hours a stone lying on the surface of a glacier will have been carried downwards and seawards two or three feet. When a glacier in an Arctic country reaches the sea, great pieces of ice break off the end and float off into the sea. These pieces of ice are called ice-bergs. Some of these ice-bergs are very large, perhaps several hundred feet high, and a mile or more in length. But when an ice-berg floats in the sea, about seven-eighths of its mass is above the surface, for the ice is partly buoyed up by bubbles of air which are scattered throughout the ice-bergs.

If two pieces of ice be gently pressed together, they stick fast, and cannot be separated without difficulty. We may often see, on the surface of a frozen pond, pieces of ice which have stuck fast and become frozen to the surface. A special name has been given to this freezing together of pieces of ice. It is called **Regelation.** If, however, the pieces of ice be very cold they will not freeze together by regelation. To allow of their sticking fast their surfaces must be melting.

One of the most curious and interesting experiments with ice is one that requires a good-sized piece of ice. If we can get a fairly big block, we may support it on the edges of two tables, or two chairs, placed near each other. When it is thus supported we may hang over it a piece of wire, to the ends of which heavy weights are attached. After a little while we shall see that the wire has begun to eat its way into the ice. Soon it will have eaten half way through, and ere long the weights will fall to the ground. The wire will have passed right through the block of ice, and only a hazy streak will show where it cut its way through, for the block of ice will be as solid as ever. Let us try and understand how the wire is able to pass through the ice. We have learnt that when water is converted into ice it expands; when ice is converted into water it of course contracts or diminishes in volume. Now suppose we take a block of ice, and by great pressure squeeze it into a smaller volume, we squeeze it into water, or in other words we cause the ice to melt. This is just what takes place when the weighted wire presses upon the block of ice. The pressure of the wire causes the ice immediately beneath it to melt, and then the water formed by this melting collects on the top of the wire. But when it is above the wire it is no longer pressed by the wire, and it therefore expands again and freezes. This goes on throughout the whole mass of the ice; the ice being squeezed into water beneath the wire, rising, and then expanding into ice again above the wire.

Note. A freezing mixture may be made by mixing five parts of sal-ammoniac with five parts of nitre and five parts of sodium sulphate, and pouring upon them sixteen parts of cold water. If a vessel containing water be placed in this mixture, the water it contains may be converted into ice for the purposes of experiment.

XXII. BOILING WATER.

If we hold a thermometer in the steam that freely rises from boiling water, we shall see that the mercury stands at 100°. This is the temperature at which water boils at the sea level. But if we were to make this simple experiment on the top of a high mountain, we should find that the mercury would not stand at 100°, but somewhat lower. The higher the mountain the lower the temperature at which water boils on its summit. And if we were to boil water at the bottom of a deep tin mine, we should find that the mercury of a thermometer, held in the steam, would stand somewhat above 100°. The deeper the mine the higher the temperature at which the water boils at its bottom. What can be the cause of this change of boiling point? Before we answer this question let us make a simple experiment. Let us boil some water in a small test tube, and, while it is boiling, close the mouth with a tightly-fitting cork, through which a thermometer passes, the bulb of the thermometer being surrounded by the steam which rises from the boiling water. As we watch the mercury in the thermometer it rises above 100°. We must be careful only to go on heating the water long enough just to make sure of the fact that the steam from the boiling water has a temperature above 100°. Why does the water in this experiment boil at a temperature above 100°, for the temperature of the steam shows the temperature at which the water boils, the temperature of the boiling water and that of the steam being the same? Because the cork prevents the steam in the test tube from escaping, and the steam, unable to escape, presses upon the surface of the water, and tends to prevent more water-vapour particles from rising from that surface. The heat of the lamp-flame, therefore, being unable wholly to expend itself in converting more water into steam, partly expends itself in making the water and the steam above it hotter. We must not go on heating lest the pressure become great enough to burst the tube. From this experiment we see that increased pressure causes the water to boil at a higher temperature. And we may fairly infer that decreased pressure would allow the water to boil at a lower temperature. Now, we can understand why water boils at a higher temperature at the bottom of a mine, and at a lower temperature on a mountain top. At the top of a mountain the atmospheric pressure is less than it is at the sea level, while at the bottom of a mine the pressure is greater. The change of boiling point is due to change of atmospheric pressure. Let us now make another experiment. Let us boil some water in a glass flask, and, while it is boiling, tightly cork the flask, and then removing the lamp, turn the flask topsy-turvy, and place the mouth, with the cork in, just under water in a glass tumbler. If now we pour some cold water over the inverted flask, we shall see that the water begins to boil vigorously, and for a little while. The more we cool the flask the more vigorously does the water boil. It is well, however, not to carry on this part of the experiment too long. We shall probably find some difficulty in getting out the cork; and it may be necessary to place the flask, after drying the outside, over the flame again for a minute or two. We shall then find that the cork can be pulled out without difficulty. Let us see what takes place in this experiment. When the flask containing the boiling water is inverted, the water still remains at about 100°. The space above the water is occupied by water-vapour, which exerts a sufficient pressure on the surface of the hot water to prevent it from boiling. But when we pour cold water on the flask, some of the steam or water-vapour is condensed and forms liquid water-drops on the inside of the glass of the vessel. But owing to the condensation of this portion of the steam, there is not enough steam left in the upper part of the vessel to exert the full atmospheric pressure on the hot water. The pressure, therefore, is diminished, and the heat contained in the water is sufficient to make it boil at this diminished pressure. As we pour more cold water on to the vessel more steam is condensed into water: the pressure is still further diminished, and the water boils again.

There is one more point to be noticed. The temperature, 100° C., is the boiling point of the substance water. If the water contain certain other substances, or liquids, the boiling point may change. Let us test this by experiment. Let us add a little spirits of wine to some water in a flask. On holding the thermometer in the vapour that rises when this mixture boils, we shall find that the temperature is below 100°. To the water in another flask let us add some concentrated sulphuric acid. We shall see that the mercury of a thermometer held in the vapour that rises from this mixture, when it is boiled, shows a temperature of more than 100°. Thus we see that 100° is the temperature at which **pure** water boils.

XXIII. MORE EXPERIMENTS WITH BOILING WATER.

We have two or three more experiments to make with boiling water. Let us take a barometer tube, and filling it quite full of mercury invert it over a basin of mercury. There will be an empty space—the torricellian vacuum—above the mercury. Now let a little water be introduced into the tube, through the open end, beneath the mercury in the basin. Being lighter than the mercury it will rise to the upper end of the tube. There it will be in the vacuum, and vapour will at once rise from its surface.

into the vacuum. This vapour will exert a certain amount of pressure, and will depress the mercury. The mercury in the tube will sink a little. If now the water be heated, more vapour will rise into the tube; a greater pressure will be exerted, and the mercury will sink further. The more we heat the water, the further will the mercury be depressed. The power which the vapour of water possesses, of forcing down the mercury against the atmospheric pressure which tends to keep it up in the tube, is called the **tension of aqueous vapour.** We see by this experiment that the more the water is heated, the greater is the tension of the aqueous vapour. Our next experiment will tell us something more about the tension of the vapour. We must take a tube shaped like a very lop-sided U, that is, a U in which one side is a foot long, while the other side is only three inches long, thus—L. The short side is closed: the long side (or limb) is open. If we pour in a little mercury we shall, with a little care, be able to fill the short closed limb with the liquid metal, and to have only about half an inch of mercury in the longer limb. Now let us pour in, at the open end, a very little water. On turning the tube into a certain position we shall be able to cause most of this water to rise into the shorter limb, where it will occupy about a quarter of an inch at the top. The longer limb may now be fitted into a large flat cork through which it passes. Let us, in the mean time, take a large-mouthed flask and boil some water in it. While it is boiling we may put the lower portion of our bent U shaped tube into the steam, letting it be supported by the flat cork, which will rest lightly on the mouth of the tube. After a little while we shall see the water in the short closed limb begin to give off vapour of water. This forces down the mercury in that limb, until the level of the mercury is the same in both limbs. If we leave it a minute or two we shall probably find that the mercury stands a little higher in the long limb than in the short limb. If so we must take the tube out of the steam. When the water-vapour cools in the short tube, it will turn again to water, and the mercury will rise again in the short limb. But we shall see that there is a little bubble of air at the top of the water. Water dissolves a certain amount of air, as we have already seen. But when the water is boiled this air is forced out of the water. This bubble of air was originally dissolved in the water, and has been forced out by the heat. We shall not find much difficulty in getting rid of it by inclining the tube and letting it escape into the longer open limb. After we have got rid of this bubble of air, which was interfering with the accuracy of our experiment, we may put the U shaped tube back into the flask in which water is being boiled. When the water in the short tube is again raised to 100°—the temperature of the steam rising from the boiling water in the flask—the mercury will again be depressed in this tube, and we shall see that the level of the mercury in the two limbs of the tube—the long limb and the short limb—is **exactly the same**. Now what does this experiment teach us? The water in the short tube is at the boiling point, for it is immersed in the steam which rises from the boiling water in the flask. The mercury in the short tube is depressed because of the tension of the vapour of water at the boiling point. Let us consider what is the amount of the pressure due to the vapour of water in the short tube. Surely the fact that it presses down the mercury until the level of the mercury is **the same** in both limbs of the tube, shows that the pressure due to the vapour of water in the short tube, is exactly equal to the pressure on the surface of the mercury in the long tube. And since the long tube is open to the air the pressure on the surface of the mercury in this tube is the atmospheric pressure. The pressure of the vapour of the boiling water in the short tube is therefore equal to the atmospheric pressure in the long tube. Hence it is said that water is at the boiling point when the tension of its vapour is equal to the atmospheric pressure. Let us, before putting away our U shaped tube, make one more experiment with it. Let us take it out of the steam, and when it has got cool, let us add just a drop or two of spirits of wine to the water, and let this water and spirits rise to the top of the shorter limb, instead of having pure water there. On putting it back into the flask in which the water is boiling, we shall soon see that the mercury is depressed in the short tube; but instead of stopping when the level is the same in both tubes, it goes on sinking until all the mercury is driven out of the short tube, and bubbles of vapour make their escape, rising through the mercury in the long tube. What does this show? Surely this: that the tension of the vapour given off by the mixed water and spirits of wine, when this mixture is raised to a temperature of 100°, is **greater than** the atmospheric pressure. It is only when we use pure water (and at the sea level), that the tension of its vapour at 100° is equal to the atmospheric pressure.

XXIV. LATENT HEAT.

WATER exists in three states; as a solid in ice; as a liquid in ordinary water, and as a gas in aqueous vapour. In this substance changes of state are, under ordinary conditions of atmospheric pressure, brought about by changes of temperature. But besides changes of state, water also undergoes changes of volume, and these changes of volume are also due to changes of temperature. Let us in this lesson learn what we can about the changes of volume and the changes of state which result when the substance water is gradually heated.

Let us start with a low temperature, and follow in thought an imaginary experiment. Suppose, then, that we have a pound of ice at a temperature of −10° C. and that we are able exactly to measure its size. On raising the temperature there is a slight expansion of the ice. This expansion continues until we reach a temperature of 0° C. Then occurs a change of state; the ice melts. We go on heating until all the ice is melted, and we find that the temperature is still 0° C. The ice has been converted into ice-cold water, and the volume has diminished. The reduction of volume is about one-twelfth. On continuing the application of heat, the temperature rises and the volume continues to decrease, until a temperature of 4° C is reached. At this temperature water is in its heaviest and most condensed state; and this temperature, 4°, is, therefore, called the point of **maximum density** of water. As the temperature of the water rises above 4° the water begins to expand, and the expansion continues up to a temperature of 100° C. At this temperature the second great change of state occurs; the water boils. We go on heating until all the water has boiled away, and we find the temperature of the water-vapour produced is still 100° C. An enormous increase of volume has occurred; such that every cubic inch of the water has become nearly 1700 cubic inches of the water-vapour. On continuing the application of heat the temperature again begins to rise, the water-vapour continuing to expand with the gradual increase of temperature.

Let us condense these facts into a tabular statement.

1. From –10° to 0° the ice expands.
2. At 0° the ice melts: much contraction.
3. From 0° to 4° the water contracts.
4. From 4° to 100° the water expands.
5. At 100° the water boils: great expansion.
6. From 100° to 200° the water-vapour expands.

Some of these facts may be illustrated by experiment.

1. The expansion of ice between –10° and 0° is merely a particular example of the general fact that solids expand on the application of heat. It is not very easy to exhibit the expansion of ice; but it is quite easy to exhibit the expansion of another solid. I have here a metal ball which, when it is cold, will easily pass through this metal ring. I heat the ball over the spirit lamp. It expands with the heat, and now it will not pass through the ring. Gradually it cools; and as it cools it contracts; soon it has contracted so much that it falls through the ring.

2. That water expands when it passes into ice (from which it may be inferred that ice contracts when it passes into water) has already been illustrated at the beginning of Lesson XXI.

3 and 4. Take a tall glass jar three parts full of water at the ordinary temperature. Put in some pieces of ice at the top. The ice cools the water, and melting itself becomes water at zero. As the ice melts, therefore, water at various degrees of temperature down to zero is produced. The heaviest water will of course collect at the bottom. We shall find, by means of our thermometer, that this heaviest water is at a temperature of 4°. A little care will enable us to obtain either colder or warmer water above this heaviest layer at 4°.

5. That great expansion takes place when water boils, may be shown by lightly corking a strong test tube, and boiling a little water at the bottom. The cork will be forced out of the tube.

6. The fact that water-vapour expands between the temperatures of 100° and 200°, is again merely a particular example of a general fact, that gases expand on the application of heat, a fact that has already been shown in Lesson XIII.

We have now to take into consideration a very important point. As long as we heat the ice from -10° to 0°, the heat we apply warms or raises the temperature of the ice; but after this, as long as the ice is melting, the heat we apply converts the ice into water, but **leaves the temperature unchanged.** So too, as long as we heat the water from 0° to 100°, the heat we apply warms or raises the temperature of the water; but after this, as long as the water is boiling away, the heat we apply converts the water into water-vapour, but **leaves the temperature unchanged.** From this we see that a certain amount of heat must pass into, and be absorbed by, the ice before it can become water; and a certain amount of heat must enter into, and be absorbed by, the water before it can become water-vapour. Whenever ice passes into water, heat is absorbed; whenever water passes into water-vapour, heat is absorbed; but whenever water-vapour passes into water, heat is given out; and whenever water passes into ice, heat is given out. It thus appears as if heat were hidden away in water-vapour, and only made its appearance when the water-vapour becomes liquid water; and as if that were hidden away in water and only made its appearance when the water becomes solid ice. Hence this heat which is thus hidden away is called **Latent Heat.**

We may now make one or two experiments. Place two beakers side by side over two equal spirit flames; and let there be some ice in one beaker, and about the same weight of ice-cold water in the other beaker. When all the ice is just melted place the thermometer first in one vessel then in the other. The water in one beaker (that which held the ice) will be at 0° or a little higher: the water in the other will be at 60° or 70°. The water in one will feel ice-cold, that in the other will be quite hot. Instead of heating both beakers we may only heat that one which contains the water, and, when the temperature is about 80°, we may pour this hot water into the beaker containing the ice. All the ice will be melted, but the temperature, after it is melted, will be very little above zero. If we could perform an experiment of this sort quite accurately, taking equal amounts of water and ice, we should find that the water must be at a temperature of 79° C., just to melt all the ice.

The fact that ice absorbs heat in becoming water is only a particular example of the general fact that, whenever a solid is liquefied, whether by melting or by solution, a certain amount of heat is rendered latent. Place a few crystals of nitre on the palm of your hand, and pour a drop or two of water upon them. They dissolve; heat is rendered latent; and your hand is cooled by the abstraction of some of the heat by the liquefying nitre. Put some of the nitre in the bottom of a test tube, and pour some water, the temperature of which you have ascertained, upon it. The nitre melts, and the temperature of the water falls several degrees.

So too the fact that water absorbs heat in becoming water-vapour, is only a particular example of the general fact that, whenever a liquid is converted into vapour, heat is rendered latent. Pour a few drops of ether on your hand; it will rapidly evaporate, or pass into vapour. Heat is rendered latent, and your hand is cooled by the abstraction of some of its heat by the vapour. Tie a piece of wet rag round the bulb of your thermometer. The mercury will show that the bulb is cooled by the evaporation of water from the wet rag.

We will make one more experiment. Take some solid phosphorus pentoxide (the white smoke formed when phosphorus is burnt in the air or in oxygen) and pour some water upon it, having previously ascertained the temperature of the water. Instead of being cooled the water will rise considerably in temperature. Why is this? Because something more than mere solution has taken place. The phosphorus pentoxide **unites chemically** with the water, and chemical union gives rise, as we have before learnt, to heat. In this case the heat produced by chemical union is much greater than the loss of heat due to the solution of the phosphorus pentoxide. The heat produced on dissolving phosphorus pentoxide is, therefore, only an **apparent** exception—not a real exception—to the rule, that, whenever a solid is converted into a liquid, a certain amount of heat is rendered latent. A certain amount of heat is rendered latent in this case also; but far more heat is set free by the chemical action which takes place.

To return to the substance with which we started, water. Whenever water becomes water-vapour, heat is rendered latent or is hidden away in that water-vapour; and whenever that water-vapour is condensed (as cloud, mist, or rain) into liquid water, the latent heat is set free. So too, whenever ice melts, heat is rendered latent or is hidden away in the water produced; and whenever that water freezes again, the latent heat is set free.

XXV. A LESSON ON HEAT.

THE handle of the teaspoon with which we stir our hot tea becomes hot; and some of us may have noticed that a solid silver spoon becomes hotter than one that is electro-plated. The handle of the cup which contains our tea does not become nearly so hot. These are matters of everyday observation. They illustrate, however, that which is called the **conduction** of heat; and they show that some substances conduct heat better than others. Here is a piece of glass tube about five inches long. I can hold it at both ends while I heat the middle in the lamp; the middle soon becomes red-hot and softens, so that I can bend the tube, or pull it apart. But the ends are not so hot as to be uncomfortable to my fingers. Here is a piece of thick copper wire also about five inches long. I heat that, too, in the middle, while I hold it at both ends. But very soon the ends become too hot to hold. In the case of the wire, the particles in the middle rapidly hand on the heat they acquire from the flame, to their neighbour particles, and so on to the very ends of the wire. The wire conducts heat well. In the case of the glass the particles do not hand on the heat to their neighbours, they keep it all for themselves, and so become red-hot. The glass does not conduct heat well. Let us now take a small test tube and fill it nearly full of water. If we hold it by the bottom in an inclined position, we can, by means of the lamp, heat the water near the top until it boils: the water at the bottom will, however, be still quite cool. This shows that water does not conduct heat well. If instead of water we put mercury in our test tube, and heat the liquid metal near the top of the tube, we shall soon find that the bottom of the tube becomes too hot to hold. The mercury conducts heat so well that the quicksilver at the top rapidly hands over to the rest of the metal the heat that it has acquired from the lamp. If, however, we heat our test tube of water at the bottom, we shall soon get the whole of the water to boil. How is it that the water conducts the heat so much better in this case? Let us consider what happens. The water near the bottom soon gets hot; but water when it is heated expands and becomes lighter. The hot water therefore rises, and cold water comes in to take its place. This water also gets heated and rises. Thus a **circulation** of the water is set up, and the heat is distributed throughout the water by this process of circulation. If instead of a test tube we take a small beaker, and heat it over the lamp, and then very carefully pour in, at one side, some litmus solution, we shall be able to see the circulation. The coloured liquid is carried downwards by the cold current flowing down the sides, and is then carried upwards by the hot current ascending in the middle. When a mass of liquid is heated in this manner, the heat is said to be distributed by **convection.**

Let us notice, in passing, that the action of the sun on the ocean is like the action of the lamp in our first experiment with water, where the heat was applied near the surface. The sun beats down on the surface of the ocean, and heats the water there: but the heated water, being light, stays at the top. Under the equator, however, the surface water of the Atlantic is swept westwards by the prevailing easterly winds (N. E. in the northern hemisphere, S. E. in the southern hemisphere). These winds are called trade winds. By them the water is swept westward into the Gulf of Mexico, whence the hot water flows out, through the Florida Straits, as the Gulf Stream.

In air and in other gases, heat is distributed very much more readily by convection than by conduction. If we take a glass flask, and put a small piece of phosphorus at the bottom, and then light the phosphorus (by gently heating the bottom of the flask over the flame), the heat of the burning phosphorus will set up a circulation of the air, which will be made visible by the white smoke of phosphorus pentoxide, produced by the combustion. The hot air rises in the middle, and cold currents flow down the sides, and the smoke particles enable us to watch the circulation of the air. A circulation of this kind goes on on a large scale in the atmosphere. The land and sea under the equator become heated; and the heat is communicated to the air that rests on the land or sea. The hot air rises, and colder currents of air blow in, from north and south, to fill up the spaces left by the ascending hot air. And just as, in the experiment, the circulation is completed at the top of the flask, so, on a large scale in nature, the circulation is completed in the upper regions of the atmosphere. The winds which blow in towards the equator are the north-east and south-east trades. They would be north and south winds, if the earth were not rotating on its axis. If, while you are drawing a straight line on a piece of paper, someone pulls the paper away to one side, your line is no longer straight. The motion of the paper from under your hand has given it a slanting direction. In somewhat the same way the earth slips away from under the wind which, so to speak, starts with the intention of going in a straight line north or south; but the rotation, or slipping away of the earth from under it, gives it a slanting direction.

XXVI. ANOTHER LESSON ON HEAT.

We have seen that heat may be conducted from particle to particle of a material substance: and we have seen that heat may be distributed by the process of convection, a circulation being set up among the particles of a liquid or gaseous substance. But the heat that reaches us from the sun comes neither by conduction nor convection. It is said to be radiated from the sun. It comes to us in waves. We are well acquainted with the waves of the sea; and sound is carried to our ears by waves of air. But there is no water and no air to bring us the waves of heat from the sun. What then brings us these waves? It is believed that the space between the sun and the earth is filled with something called the **æther**, and that waves of heat, communicated by the sun to this æther, are by it transmitted to the earth. All hot bodies radiate away their heat, and continue to do so as long as they are hotter than the air and other things round them. But some radiate away their heat more rapidly than others. Metals radiate their heat slowly, and that is why the tea keeps hot longer in a bright metal tea-pot than in one made of earthenware. Rocks and stones are good radiators: and they readily absorb the heat radiated from the sun. It is, in fact, always true, that if a body is a good radiator, it is also a good absorber. The rugged rocks on high mountain summits get very hot during the day, but they radiate their heat as rapidly at night, especially when the air is dry; for it is found that when there is much aqueous vapour in the air, the radiation of heat is very much checked.

We have already learnt that the temperature of any substance is measured in degrees by means of a thermometer. This does not tell us, however, the amount of heat contained in any substance. Perhaps it may be said that surely, if we know the temperature of any substance, that temperature shows the amount of heat it contains. But suppose we have a test tube full of water, and a kettle full of water; the temperature of the water in each case may be 80°. But it is clear that the amount of heat in the kettleful of water at 80°, is greater than the amount of heat in the test tubeful of water at 80°. From this illustration we see that there is at least one thing—the amount of the material—that has to be taken into consideration, in dealing with amounts of heat, but has not to be taken into consideration in dealing with temperature. Let us, however, suppose this difficulty got rid of; let us suppose that we always speak of equal weights in dealing with amounts of heat. Even then the temperature of that weight of substance will not tell us what is the amount of heat contained in it. For the same weights of different substances at the same temperature contain different amounts of heat. In other words, the same weights of different substances require different amounts of heat to raise them to the same temperature. This fact is shortly expressed by saying that different substances have different **specific heats.** Let us try an experiment. Let us take the same weight of water and of mercury in two different test tubes. From its great weight the volume of the mercury will be less than one-thirteenth of that of the water. Let us start with them at the same temperature, and then let us hold the test tube containing the water for one minute (by the watch) in the flame of the lamp, and observe, by means of our thermometer, how many degrees it gains in this time. After noting this down, let us hold the test tube containing the mercury for exactly the same time and, as far as possible, in the same manner in the flame. On taking the temperature with the thermometer, we shall find that the number of degrees gained by the mercury is very much greater than the number of degrees gained by the water. Even if we take equal volumes instead of equal weights, we shall find that the mercury has gained nearly twice as many degrees as the water. Very careful experiments have, in fact, shown that a given weight of mercury will gain 30° for every 1° gained by the water, the same amount of heat being put into each. To measure amounts of heat, therefore, we cannot make use of the degrees marked on our thermometer unless we state, first, the substance, and, secondly, the weight of the substance employed. Amounts of heat are measured in **thermal units**; a thermal unit being the amount of heat necessary to raise **one pound** of **water** 10° C. From what we have learnt above we may see that one thermal unit will raise one pound of mercury 30°, or, what comes to the same thing, thirty pounds of mercury 1°. The specific heat of any substance is the amount of heat, in thermal units, required to raise one pound of that substance 1° C. in temperature. The specific heat of water is 1; that of mercury is therefore $\frac{1}{30}$.

We have now learnt how heat is conducted, distributed by convection, radiated, and measured. One question remains. What is heat? A simple answer, without much illustration, must suffice. Heat is the very rapid vibration of the molecules of a substance; a molecular tremor in the substance. That tremor may be brought about in various ways. It may be produced in the particles of this piece of lead by simple hammering. The hammered lead becomes hot; the rapid shocks of the hammer having set the lead particles vibrating. Or the tremor may be produced in the lead particles by putting the piece of lead in the sunshine. The heat waves of the æther influencing the lead molecules, set them vibrating with the tremor of heat. In a candle flame, the heat is due to the clashing together of the oxygen of the air with the hydrogen and carbon of the hydro-carbon vapour. Owing to the vigour of the chemical union, the new molecules of carbonic acid gas and water-vapour vibrate rapidly with that peculiar tremor which we call heat, and which gives rise in us to feelings of warmth.

XXVII. A LESSON ON LIGHT.

THE same æther which brings us waves of heat from the sun, brings us also waves of light. The rapidity with which these waves travel is almost inconceivable. The sun is about ninety-two millions of miles distant from the earth, and yet the light reaches us in about eight minutes. The waves of light and heat travel at a rate of nearly 186,000 miles in a second. It may be asked, How can this be found out? A few words of explanation may be given. Galileo, the great Florentine astronomer, was the first to discover that the planet Jupiter has satellites which revolve round him, just as the moon revolves round the earth. At certain times these satellites pass behind Jupiter, and then become invisible from the earth. They are then said to be eclipsed. It is well known how accurate the predictions of astronomy are. But it was found that the predictions of the eclipse of Jupiter's moons were not always quite accurate. Sometimes the predictions were right, sometimes they were a little wrong. Roemer, a Danish astronomer, showed how the error crept in; and he made this error the means of calculating the velocity of light. The earth sweeps round the sun in a nearly circular path. Now Jupiter is much farther from the sun than the earth is: and so, as the earth sweeps round, it is sometimes nearer to Jupiter than it is at other times. The distance between the two planets, in fact, varies by the whole breadth of the nearly circular path or orbit of the earth—by a distance, that is, of about 184,000,000 miles. Roemer, therefore made observations of the eclipse of one of Jupiter's satellites when the earth was nearest to Jupiter, and then made calculations at what time it should take place on a subsequent occasion, when the earth was farthest from Jupiter. The eclipse did not take place at this time, but 16½ minutes later. Why? Because the light took 16½ minutes to cross the space of 184,000,000 miles. It must have travelled, therefore, at the rate of 186,000 miles a second. There are other ways in which the velocity of light can be measured; but these cannot be described here.

Let us now make a few simple experiments with the light of a common candle. Let us first notice that the light travels in straight lines. This is so obvious as scarcely to need an experiment. But if we take a pea-shooter and look through it at the candle, we shall see that the light comes straight down the tube. And if our pea-shooter be crooked, we shall not be able to see the light of the candle. If we hold the candle before a looking-glass, we shall see a **reflection** of the candle in the looking-glass. We see no reflection of the candle, however, on the surface of a piece of white paper. In the case of the looking-glass, the light rays from the various parts of the candle-flame fall in straight lines on the mirror, and from the mirror the light from each of the various parts of the flame is reflected, true and unerring, in straight lines to the eye. But in the case of the paper, though the light rays from the various parts of the candle flame fall in straight lines on the white surface, they are scattered away from the surface in all directions,—still of course in straight lines, but no longer keeping their relative positions; so that the paper reflects only a confused blur of light. Let us now prick a largish pin-hole in a piece of paper, and hold the paper near the candle, so that we can see the light of the flame through the pin-hole. If now we hold a second piece of thin white paper about an inch from the first piece, and between it and our eye, we shall see a bright image of the candle—but the image will be inverted, or topsy-turvy. This is because the rays of light from the candle-flame cross each other as they pass through the pin-hole. Each part of the candle sends through the pin-hole its own particular tiny ray, and since the pin-hole is small, there is no confusion of the rays, but each falls into its place, and helps to form the inverted image of the flame. We could not have a better illustration of the fact that the light proceeds in straight lines. Have you ever noticed that the light which struggles through the leaves of a tree always makes a number of **round** patches of light on the ground? If so, you may perhaps have wondered why all the patches were so round. Our simple experiment will help you to understand this matter of common observation. Each round patch is an image of the sun, formed just in the same way as the image of the candle is formed in the experiment. But, instead of the pin-hole in the paper, we have the chinks between the leaves of the tree.

Leaving our candle-flame for the present, let us make one more experiment in the following way. Let us take a tea-cup, and place a shilling at the bottom. Let us now stand in such a position that the shilling is just hidden from our view by the edge of the cup. Now pour in some water. As the water rises in the cup, the shilling comes into view. It has not moved; but the presence of the water has somehow enabled us to see the shilling that was invisible before. The fact is that, though under ordinary circumstances, in air, the rays of light travel in straight lines, when rays of light pass out of water into air, or out of air into water, they make a bend just at the surface of the water. When we pour water into the cup, the rays of light coming from the shilling make this bend at the surface of the water, and thus pass over the edge of the cup and reach our eye. This bending of light at the surface of water (and other transparent liquids) is called **refraction**. The curious thing is that we have grown so accustomed to seeing things in straight lines from our eye, that we fancy the fish we see at the bottom of a river is in a straight line from our eye. But it is not, because of the refraction. A rifle bullet shot straight at the fish (supposing that it kept an unerring straight course) would never hit the fish, because the fish is not really there, but only seems as if he were there. Fill a basin pretty full of water and then drop in a threepenny bit. Look at it with the eye just over the edge of the basin, and try by a rapid plunge to put your finger on it. If you aim truly at the image, you are sure to miss the coin.

XXVIII. ANOTHER LESSON ON LIGHT.

THE waves of the æther that, when they fall upon our eyes, give rise to the sensation of white light, are not all of the same size, the larger waves being nearly half as large again as the smaller. Now if we cause the light of the candle flame to pass through a triangular-shaped piece of glass, called a prism, we shall find that the light which passes through the glass is not only refracted but is no longer white. The white rays of the candle flame have, after passing through the glass prism, assumed all the colours of the rainbow. What then has the passage through the glass prism done to the light to cause so great a change? It has sorted the waves according to their size. And when the waves are thus sorted, the sensation produced is no longer that of white light, but of a series of rainbow colours. Let us place our candle in a tolerably deep box with a vertical slit in it, through which a line of light may pass and fall upon the wall. Or if it be daylight and the window have shutters, let us close the shutters, allowing only a vertical line of light to pass through a crack or chink. Now let us hold the glass prism in such a way that the line of light must pass through it before it reaches the wall. We shall find that the vertical line of white light has become a coloured band, the colours being red, orange, yellow, green, blue, indigo, and violet. Each of these shades into the colours on either side of it. Here we have a band of sorted light-waves, the longest being in the extreme red, the shortest in the extreme violet. If we hold a stick of red sealing wax in this coloured band,—which is called a **spectrum**,—we shall see that it has its usual colour in the red rays, but elsewhere it is nearly or quite black. Why is it black? Let us consider first what takes place—so far as the light waves are concerned—when white light falls on any coloured object. When white light falls on a stick of red sealing wax, all the waves, except the large red waves, are absorbed or lost in the sealing wax. They disappear altogether, and the sealing wax passes on the light, robbed of all its waves except those which are red. In the same way this blue flower robs the white light it receives of all waves except the blue, and these only it passes on to our eyes. The white paper does not absorb any of the waves; while the black letters stamped upon it absorb the waves of all sizes.

To return to the sealing wax held in the coloured spectrum, we may now see why it appears black in all the colours except the red. It is because it absorbs all waves other than the red. It can only pass on to our eyes the red waves; and at the violet ends of the spectrum it receives no red rays.

We have seen that the red waves are nearly half as big again as the violet waves; and yet these red waves are so minute that there are something like 39,000 in a single inch. There are also waves that are somewhat bigger than these red waves; but they excite no sensations when they are received on to the sensitive part—or **retina**—of the eye. They are not light-bearing waves; but they are heat-bearing waves. The waves of light and of heat are thus intimately associated.

If we are reading a printed page by means of our candle, we know that the nearer we are to the candle flame the more light falls on our page. If we move farther away less light falls on the page. We might suppose at first that if we doubled the distance from the candle, the light falling on our page would be halved: and if we trebled the distance, the light would be one-third of what it was. But this is not so: the light diminishes more rapidly than this. At double the distance, the light which falls on our page is reduced to one-fourth. If we treble the distance, the light is reduced to one-ninth. While at four times the distance, the page is only illuminated with one-sixteenth of the previous amount of light. Perhaps some may understand me when I say that, as the page is farther and farther removed, the light is reduced to a fraction, represented by one over the square of the distance.

We may now, in conclusion, consider how the intensity of any given source of light may be measured. Suppose, for example, we wish to know how much brighter an oil lamp is than one candle flame. One way is as follows: Hold a stick before a white wall in such a way that two clearly defined shadows are cast on the wall, one from the lamp, the other from the candle. That cast by the lamp will probably be the darker. Now move the lamp away from the stick until the two shadows are equally dark. When this is the case, the light falling on the wall from each source, candle and lamp, is of equal intensity. If then we measure the distances, and remember in what proportion the light decreases as the distance increases, we can tell how much brighter the lamp is than the candle flame. If, for example, the lamp is twice as distant as the candle, its light is four times as intense. If it is two and a half times as distant, its illuminating power is six and a quarter times as great. Another simple way of comparing the relative intensity of the two lights, is to place them some distance apart, and to place between them a piece of white paper, in the middle of which is a grease spot. Our object is to find the position in which the paper is equally illuminated on both sides. When it is not equally illuminated the grease spot appears dark when we look at it from one side, and bright when we look at it from the other side. The paper must be moved backwards and forwards until the grease spot disappears from view. The paper is then equally illuminated on both sides, and we may measure the distances of the two sources of light. The illuminating powers are, as before, proportional to the squares of the distances from the paper. Such contrivance for measuring the illuminating power of different sources of light is called a **photometer.**

XXIX. A LESSON ON WEIGHT.

One of our earliest and most familiar experiences is, that all the various substances which we come across in our everyday life have weight, and that if we let any solid or liquid substance drop from our hands, it falls at once to the ground. And it falls straight downwards, without deviation to the right hand or the left. The mason who is building a wall takes advantage of this fact. He hangs a heavy weight to the end of a piece of string. The string prevents the weight from reaching the ground, but becomes stretched in the direction towards which the heavy body falls. And since the weight tends to fall straight downwards the string is stretched in this direction. The plummet gives us a vertical line, just as the surface of a still liquid gives us a horizontal line. There are, however, apparent exceptions to the rule that all substances fall to the earth, and that they fall straight downwards. A balloon filled with hydrogen does not fall at all, but ascends in the air. Yes; but so does a cork ascend in water. And why? because the water buoys up the cork. And so too, in just the same manner, does the air buoy up the balloon filled with hydrogen. And if we drop a flat piece of thin paper it does not fall straight downwards. But here again it is the air that is to blame. It is merely because the piece of paper is flat and thin, and so exposes a large resisting surface to the air, that it floats from side to side as it falls. If we crumple it up into a ball it will fall nearly, if not quite, straight downwards. In a vacuum not only do all bodies, of whatever shape size or weight, fall quite vertically, but they all—sovereigns, stones, corks or feathers—fall at the same rate.

If we let a round smooth stone fall from the top of a cliff, or down a well, we can, by noting the time which the stone takes in reaching the bottom, find out what the height or depth is. Such a stone will fall about sixteen feet in the first second. In the next second it will fall forty-eight feet, that is to say, 16 + 32 feet. In the third second it will fall eighty feet, that is to say, 48 + 32 feet. In the fourth second it will fall 112 feet, that is to say 80 + 32 feet. In the fifth second it will fall 144 feet, that is to say 112 + 32 feet. And in succeeding seconds it will always, in any one of them, fall 32 feet more than it fell in the one immediately before. Suppose a stone be dropped down a deep well and reach the bottom in three seconds. It will have fallen 16 feet in the first second, 48 in the next, and 80 in the last; in all 144 feet. That which makes the stone fall is called the force of **gravitation.** We must notice that, under the influence of the constantly acting force, the stone falls with constantly accelerated speed.

The force of gravitation, which gives rise to what we call weight in material substances, is not confined to bodies at or near the surface of the earth. It is universal in its action. It is this force that binds the earth to the sun, and the moon to the earth. But its action decreases as we increase the distance from the centre of the attracting body. And it decreases just in the same way as the illuminating power of light decreases as we leave the source of light. That is, if we double the distance, the force of gravitation becomes one-fourth; if we treble the distance, it becomes one-ninth; if we increase the distance by four times, the force of attraction is reduced to one-sixteenth, and so on; the attraction being always reduced to a fraction represented by one over the number representing the distance multiplied by itself —or squared, as mathematicians call it.

The English measures of weight are the pound, ounce, grain, and so on. But men of science now nearly always use the French measures of the metric system. When once we know decimal fractions, this metric system is very simple. The unit of weight is called a **gram**; and it is equivalent to 15.43 English grains. A weight of ten grams is called a **deca**gram; 100 grams a **hecto**gram; 1000 grams a **kilo**gram; $\frac{1}{10}$ of a gram a **deci**gram; $\frac{1}{100}$ of a gram a **centi**gram; $\frac{1}{1000}$ of a gram a **milli**gram. The unit of length is called a **metre**. It is equivalent to 39.37 English inches. A length of ten metres is called a **deca**metre; 100 metres a **hecto**metre, and so on. Using the metric system, therefore, the stone dropped down a well will fall 4.88 metres during the first second, 14.64 during the next, 24.4 during the third, and so on, adding 9.76 metres for each second to the distance fallen in the previous second.

Here are the Tables of Weight and of Length.

TABLE OF WEIGHT.

10 Milligrams	make	1 centigram.
10 Centigrams	„	1 decigram.
10 Decigrams	„	1 gram.
10 Grams	„	1 decagram.
10 Decagrams	„	1 hectogram.
10 Hectograms	„	1 kilogram.

TABLE OF LENGTH.

10 Millimetres	make	1 centimetre.
10 Centimetres	,,	1 decimetre.
10 Decimetres	,,	1 metre.
10 Metres	,,	1 decametre.
10 Decametres	,,	1 hectometre.
10 Hectometres	,,	1 kilometre.

The unit of capacity is called a Litre. It is a little more than 1.75 English pints. There are exactly 1000 cubic centimetres in a litre. One cubic centimetre of pure water weighs one gram.

XXX. SPECIFIC GRAVITY.

We have a common expression, 'as heavy as lead.' Lead is not by any means, however, the heaviest metal. Even among the commoner metals, mercury and gold are heavier than lead. Platinum, a white metal sometimes used in jewellery, is heavier still. In comparing the weights of metals, water is taken as a convenient standard; the weight of a metal, or other solid or liquid body, compared with water, is called its **Specific Gravity**. The specific gravities of some of the elements are given in the form of a table at the end of this lesson.

To find the specific gravity of a solid substance that does not dissolve in water, we may proceed as follows: Weigh the substance in the air; then suspend it in water, and see what it weighs in water. It will weigh less in water than it does in air. Subtract the weight in water from the weight in air. This gives the weight of a bulk of water equal to that of the substance. Now state in proportion thus—

Wt. of equal bulk of water: wt. of subst. : : sp. gr. of water : sp. gr. of subst.

Or in an equation thus:

$$\text{Sp.gr.} = \frac{\text{wt. of substance in air.}}{\text{wt. of subst. in air} - \text{wt. of subst. in water.}}$$

For example, a piece of sulphur weighs 7.4 grains in the air. In the water it weighs 3.7 grains. The weight of an equal bulk of water is therefore $7.4 - 3.7 = 3.7$ grains: then—

$$\underset{\text{wt. of = bulk of water.}}{3\cdot7} \quad : \quad \underset{\text{wt. of sulphur.}}{7\cdot4} \quad :: \quad \underset{\text{sp. gr. water.}}{1} \quad : \quad \underset{\text{sp. gr. S.}}{2}$$

or according to the equation:

$$\text{Sp. gr. of S.} = \frac{7\cdot4}{7\cdot4 - 3\cdot7} = \frac{7\cdot4}{3\cdot7} = 2$$

We have now to inquire why the difference between the weight of a substance in air and its weight in water is equal to the weight of a bulk of water equal to that of the substance. Make a diagram of a cube (say) of glass suspended in a glass of water. Let the upper face of the cube be parallel with the surface of the water. Mark this A. B. Mark the corresponding lower face, C. D. On the surface A. B., there is a column of water extending from A. B. to the surface of the water. This is pressing the cube downwards. On C. D. there is an upward pressure equal to that of a column extending from C. D. to the surface. This may seem surprising, so we will illustrate it by experiment. Take a glass cylinder open at both ends. Hold firmly against one end (which must be quite smooth) a flat glass plate. Dip this end well under water, and let go of the flat glass plate. It will not fall to the bottom, but will be held to the open end of the glass cylinder by the **upward pressure of the water.** We have still to show that this upward pressure is equal to that of a column of water reaching to the surface. To show this we may pour water into the upper end of the cylinder. Directly the water in the cylinder reaches the level of the water outside, the glass plate drops off. This is because the downward pressure of the water in the cylinder just counterbalances the upward pressure on the under surface of the glass plate.

To return to our cube of glass. The upper surface is pressed downwards by the pressure of a column of water from A. B. to the surface. The under surface is pressed upwards by the pressure of a column of water from C. D. to the surface. This upward pressure is therefore the greater. And how much the greater? It is greater than the downward pressure by the same amount, as the column of water from C. D. to the surface, is greater than the column of water from A. B. to the surface. But the difference between these columns is a column of water from A. B. to C. D., that is to say, a column of water just the size of the cube. The cube is, therefore, pressed upwards, or rendered lighter, by an amount equal to the weight of a cube of water of the same size. In other words, the cube weighed in water is lighter than the cube weighed in air, by the weight of an equal bulk of water. Or to put it in yet another way, the difference between the weight of the cube in air and the weight of the cube in water, is the weight of an equal bulk of water.

But suppose I wish to know what is the specific gravity of this liquid, spirits of wine. I can find out by weighing a piece of sulphur, first in the air, then in water, and then in the spirits of wine. A piece of sulphur, for example, weighs 7.4 grains in air, 3.7 grains in water, and 4.07 grains in the spirits. Now we have seen that the weight in air (7.4) minus the weight in water (3.7) is the weight of an equal bulk of water (3.7). So, too, the weight in air (7.4) minus the weight of the spirits (4.07) is the weight of an equal bulk of spirits (3.33). Knowing, therefore, the weights of equal bulks of water and spirits, we may state in proportion thus—

Wt. of a certain bulk of water : wt. of = bulk of spirits : : 1 : answer; or by our equation,

$$\text{Sp. gr. of spirits} = \frac{\text{weight of spirits.}}{\text{wt. of} = \text{bulk of water.}}$$

That is, in our particular case,

$$\text{Sp. gr. of spirits} = \frac{3{\cdot}33}{3{\cdot}7}, \text{ or } .9$$

TABLE OF SPECIFIC GRAVITIES.

Platinum	21.50
Gold	19.30
Mercury	13.60
Lead	11.40
Silver	10.50
Copper	8.90
Iron	7.80
Tin	7.30
Diamond	3.40
Graphite	2.30
Phosphorus (red)	2.10
Sulphur	2.00
Charcoal	2.00
Phosphorus (yellow)	1.80
Water	1.00
Sodium	.97
Potassium	.86

XXXI. SOME CHEMICAL TERMS AND PROCESSES.

If we take a little salt at the bottom of a test tube, and pour upon it some water, the salt will, after a little while, dissolve. That which here takes place is called **solution**; and, as we have before seen, solution is accompanied by the passage of a certain amount of heat into the latent state. The liquid we obtain, which contains the common salt dissolved in it, is called a solution. If now to this solution of common salt, or sodium chloride, as it is also called, some solution of a substance called silver nitrate be added, what is termed a **chemical reaction** takes place. The silver nitrate acts upon the sodium chloride in such a way that a change of partners, so to speak, is effected, and silver chloride and sodium nitrate are formed. The sodium nitrate remains in solution, but the silver chloride separates from the liquid in white curdy flakes. The white silver chloride forms what is termed a **precipitate**, which, since it does not dissolve in the liquid, but remains for a while floating in the solution, is said to be held in **suspension**. The process by which a precipitate is formed is called **precipitation**. Now suppose we wish to separate the solid particles of silver chloride from the solution of sodium nitrate in which they are suspended; we must **filter**. Here is a piece of porous paper like blotting-paper. From it I cut a circular piece about five or six inches in diameter; this circular piece I double across, and then double again at right angles, so as to bring it to a three-cornered shape. This piece of paper I can now open out into a hollow cone, which will just fit into a glass funnel. Having wetted this, I pour in the solution and the precipitate which it holds in suspension. The solution, containing the sodium nitrate, soon passes through the porous paper, but the precipitated silver chloride remains on the filter. This process is called **filtration**, while the solution which passes through is sometimes called the **filtrate.** The filter and the precipitate may be left to dry, after which the precipitated silver chloride may readily be detached from the paper. We may notice that it is no longer white, but has turned a deep violet colour. To get the sodium nitrate out of the solution we must go through a different process. We must place the solution in a small shallow porcelain basin, and gradually heat it over the lamp. Soon the water will begin to rise in vapour. And if we continue the process long enough, all the water will be driven off as vapour, and the sodium nitrate will remain at the bottom of the basin in the solid condition. This process is called **evaporation.** In some cases we wish to obtain the water rather than the substance which the water holds in solution. For certain chemical operations, for example, we want perfectly pure water, free from the salts of lime and other substances which ordinary water holds in solution. In such cases we carry on the evaporation in a vessel called a **retort,** which is so made that the vapour which rises from the impure water is **condensed** to water in another part of the vessel, and drops into a second vessel called the **receiver.** This process is called **distillation.** The object of the process is to enable you to collect the substance converted into vapour during evaporation.

Spirits of wine, and brandy, are obtained by a process of distillation.

Let us now notice to how large an extent some of the processes just described are carried on in Nature. All over the surface of the ocean, and indeed wherever water is exposed to the open air, the process of evaporation is going on. And what becomes of all the vapour thus evaporated from the surface of the ocean? Some of it is condensed to rain above the ocean areas; but some of it is carried by the winds to distant mountains, and is there condensed into rain or snow. Here, then, we have a constant process of distillation, by means of which pure water is separated from the salt water of the ocean, and is collected in mountain streams and rivers. But as the river flows seawards it takes up into solution a larger or smaller quantity of carbonate of lime, common salt, or other such materials. And at the same time, especially after rain, it carries seawards large quantities of mud and sand in suspension. The mud and sand are dropped near the mouth of the river in **delta** or **estuary** deposits, but the matter in solution is carried onwards into the ocean. There minute foraminifera and other shell-bearing creatures cause the carbonate of lime to be precipitated in the solid form, and with it, as we have seen in Lessons III. and IV., build their shells. But no creatures can extract the common salt from the waters of the ocean. This therefore goes on increasing in amount, as rivers constantly bear seawards small quantities of this material. The ocean is a vast evaporating basin, in which, by the vaporization of some of the water, the salts are being concentrated in the water that remains. This process is being carried on still more rapidly in some salt lakes, such as the great Salt Lake of Utah. This inland sea has no outlet. All the water brought into the lake by rivers and streams is got rid of by evaporation from its surface. But the salt which is held in suspension by the water of these rivers and streams is not got rid of by evaporation. It accumulates in the lake, the waters of which are therefore exceedingly salt. We can imitate this natural process by evaporating some salt water in an evaporating basin, and, as fast as the level of the liquid in the vessel falls, from the loss of pure water by evaporation, adding a little more of the salt water. By carrying this on for some time, we shall obtain a very concentrated solution of salt in the basin.

XXXII. CHEMICAL SHORTHAND AND EQUATIONS.

THE smallest quantity of an element which can enter into or be expelled from a chemical compound is called an **atom.** The hydrogen atom is the lightest atom known. In speaking therefore of the weights of other atoms we say they are so many times as heavy as the hydrogen atom. To each element a sign or symbol has been given; and this symbol is generally the first letter (or sometimes the first letter and another) of the English or Classical name of the element. With each symbol a number is connected, this number representing the weight of the smallest quantity of the element which can enter into or be expelled from a chemical compound, in other words, the weight of the atom. At the end of Lesson xv. (p. 45), there is a list of some of the common chemical elements; their symbols and their atomic or combining weights, in round numbers, being placed opposite the elements.

To represent the molecule of a compound, the symbols of the elements which compose the compounds are placed side by side; if there be more than one atom of any element, a small figure denoting the number of atoms present is placed at the foot of the symbol of the element, and on the right hand side. Thus the molecule of water contains one atom of oxygen and two atoms of hydrogen: the molecule will therefore be represented thus—OH_2. Carbonic acid gas consists of one atom of carbon combined with two atoms of oxygen: it will therefore be represented thus—CO_2. Mercury oxide consists of one atom of mercury (Hg), with one atom of oxygen: it will therefore be represented thus—HgO. Every chemical compound can thus be represented in chemical shorthand. By means of this shorthand chemical changes can be readily and neatly expressed. When sulphuric acid is added to zinc, the zinc unites with the sulphur and oxygen, forming a new compound called zinc sulphate, and setting the hydrogen free. When such a change takes place a chemical **reaction** is said to occur, and such a reaction may be expressed by means of an **equation**. On the left hand side of the sign of equality (=) we place those compounds or elements which are taken in the first instance, putting + for "added to." On the right hand side of the equation we place those new compounds, or elements, which are formed, or set free, during the reaction: thus—

$$Zn + H_2SO_4 = ZnSO_4 + H_2.$$

This shorthand translated into words means: One atom of zinc, added to one molecule of sulphuric acid, give us one molecule of zinc sulphate, and one molecule of hydrogen (which is set free).

We have already learnt that when a candle burns, carbon and hydrogen unite with atmospheric oxygen to form carbonic acid gas and water. Here are the equations:

$$(a)\ C + O_2 = CO_2.$$

$$(b)\ 2H_2 + O_2 = 2OH_2.$$

Let us consider these equations closely. In the first place we write O_2, with a small 2 low down to the right, because the oxygen exists in the air in molecules, and each molecule consists of **two** atoms. The same is true of hydrogen. Two atoms of hydrogen form a molecule: we therefore write H_2. We do not know how many atoms of carbon there are in a molecule: we therefore simply write C. To the left of the H_2, and the OH_2, in (*b*) there is a large 2. This shows that **two** molecules of hydrogen have to be taken, and that **two** molecules of water are formed. We must notice that in (*b*) two molecules of hydrogen, each containing two atoms, are taken and combined with one molecule of oxygen, also containing two atoms, and that by the combination of these three molecules **two** molecules of water, each containing one atom of oxygen and two atoms of hydrogen, are formed. We must notice, too, that the number of atoms of each element to the left of the sign =, is the same as the number of atoms to the right of that sign.

Let us take one or two more equations. We have seen that if carbonate of lime be heated, carbonic acid gas and quicklime are formed. The quicklime is an oxide of the metal called calcium (Ca). The equation is—

$$Ca\,CO_3\text{ (heated)} = Ca\,O + CO_2.$$

Here we see that one molecule of carbonate of lime, containing one atom of calcium, one of carbon, and three of oxygen, splits up, under the influence of heat, into one molecule of calcium oxide (quicklime), containing one atom of calcium to one of oxygen, plus one molecule of carbonic acid gas. Again, we saw that the quicklime combines with water to form slaked-lime. The equation is—

$$Ca\,O + OH_2 = Ca\,H_2\,O_2.$$

One molecule of the calcium oxide combines with one of water to form one of slaked-lime.

And, once more, if to the slaked-lime (dissolved in water) carbonic acid gas be added, carbonate of lime and water are formed, thus—

$$Ca\,H_2\,O_2 + CO_2 = Ca\,CO_3 + OH_2.$$

Here one molecule of slaked-lime unites with one of carbonic acid gas to form one of carbonate of lime and one of water.

We made oxygen in two ways. The first of these was by heating red mercury oxide. The equation runs thus—

$$2\ Hg\,O\text{ (heated)} = 2\ Hg + O_2.$$

Two molecules of mercury oxide yield, when heated, two atoms of mercury and one molecule of oxygen.

The second method was by heating a mixture of potassium chlorate and manganese dioxide. Now in this case the manganese dioxide undergoes no permanent change. It is found, however, that if this material be added the gas comes off more readily and at a lower temperature. The equation for the action of heat on the potassium chlorate is—

$$2\ KClO_3\ (heated) = 2\ KCl + 3\ O_2.$$

Two molecules of potassium chlorate yield, when heated, two molecules of potassium chloride and three molecules of oxygen.

These equations will no doubt seem difficult at first. The difficulty is, however, soon overcome; and then by means of an equation we can express accurately in a single line a chemical change which could only be explained in half a page of ordinary writing.

XXXIII. SULPHUR.

SULPHUR is a yellow, brittle solid which is found in the neighbourhood of volcanoes. It readily combines, as we have already seen, with some of the metals. And just as the combination of a metal with oxygen is called an oxide, so is the combination of a metal with sulphur called a **sulphide.** Here is some iron wire. I coil it up into a spiral form and place it in this dry test tube, at the bottom of which there is already a little sulphur. On heating the test tube in the lamp, the wire is seen to glow. This glow is the visible sign of the chemical union which is taking place between the sulphur and the iron. Iron sulphide is formed. It is a brittle, black solid, and can be easily crushed to powder. Let us put this iron sulphide on one side; we shall find it useful presently. Many of the metals are found in the rocks of the earth's crust as sulphides—for example we find iron sulphide, or Iron Pyrites, lead sulphide, or Galena, zinc sulphide, or Blende, copper sulphide, or Copper Pyrites, besides some others.

Let us now place a little sulphur at the bottom of a dry test tube, and carefully warm it. We shall find that it melts and forms a thin yellow liquid. We may place a thermometer in this liquid, and we shall find that the mercury stands at 115°. So long as there are any pieces of solid sulphur unmelted in the light yellow liquid the temperature will remain at 115°. The heat of the flame is becoming latent in the molten sulphur. As soon as the sulphur is all melted, the temperature begins to rise. The liquid turns darker and darker in colour, and at the same time becomes thick, like so much treacle. When it is brownish red, and has a temperature of 230°,[1] the test tube may be turned topsy-turvy, for a moment, without any of the sulphur running out. If we still go on heating the sulphur it gets more liquid again, and at the same time turns almost black. At last, at 440° or thereabouts, it begins to boil, giving off reddish-brown vapour of sulphur. We may now let it cool again. When the temperature has fallen to 230° it is again thick like treacle. Let us pour some of this treacly sulphur into cold water. We shall find that the sulphur which has thus been suddenly cooled by coming into the cold water, is quite soft and elastic, and not at all unlike india-rubber. This soft elastic sulphur, however, soon becomes hard and brittle again. Crystals of sulphur may be obtained in two ways. I melt some sulphur in this crucible, and then allow it to cool. As soon as a solid crust of sulphur begins to form, I pour out as much as I can of the molten sulphur that remains.

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[1] If our thermometer is not marked to so great a number of degrees we must be careful to remove it as soon as all the sulphur is melted.

gas burns with a blue flame, some sulphur dioxide being produced, while some of the sulphur is deposited on the sides of the tube. The hydrogen at the same time unites with oxygen to form water. If, however, there be a free supply of air, all the sulphur burns as well as all the hydrogen.

When sulphuretted hydrogen and sulphur dioxide meet, the hydrogen of the one combines with the oxygen of the other, water being formed, while the sulphur of both is deposited. I have here a bottle fitted with a large cork, through which pass three glass tubes, two only just passing through the cork, the third reaching to the bottom of the bottle. To this last is connected a tube which passes out of the window. This is merely to carry off the waste gases from the bottle. The other tubes are connected, by means of india-rubber tubing, with the ends of two glass tubes proceeding from two corked test tubes. In one of these sulphuretted hydrogen is formed by the action of dilute sulphuric acid upon iron sulphide. In the other sulphur dioxide is generated by heating copper and sulphuric acid. The two gases meet in the bottle, in which yellowish white particles of solid sulphur are rapidly produced. Both sulphuretted hydrogen and sulphur dioxide are volcanic gases, so that it is very probable that the sulphur so commonly found in the neighbourhood of volcanoes is formed in this way.

The following equations may here be added. It is by no means necessary that the beginner should learn these equations by heart. He should, however, do his best to understand them.

1. For the burning of sulphur.

$$S + O_2 = SO_2.$$

One atom of sulphur unites during combustion with one molecule of oxygen to form one molecule of sulphur dioxide.

2. For the action of copper on sulphuric acid.

$$Cu + 2\ H_2\,SO_4 = Cu\ SO_4 + SO_2 + OH_2.$$

One atom of copper added to two molecules of sulphuric acid yield one molecule of copper sulphate, one of sulphur dioxide, and one of water.

3. For the action of sulphur on sulphuric acid.

$$S + 2\ H_2\,SO_4 = 3\ SO_2 + 2\ OH_2.$$

One atom of sulphur added to two molecules of sulphuric acid give three molecules of sulphur dioxide and two molecules of water.

4. For the action of iron sulphide on sulphuric acid.

$$Fe\ S + H_2\,SO_4 = Fe\ SO_4 + H_2\,S.$$

One molecule of iron sulphide with one of sulphuric acid give one molecule of iron sulphate and one of sulphuretted hydrogen.

5. For the action of sulphuretted hydrogen on sulphur dioxide.

$$2\ H_2\,S + SO_2 = 2\ OH_2 + 3\ S.$$

Two molecules of sulphuretted hydrogen and one of sulphur dioxide give two molecules of water and 3 atoms of sulphur.

XXXIV. PHOSPHORUS.

THE element phosphorus is not found in a free and uncombined state like the element sulphur. In all fruitful soils, however, it occurs in small quantities, in combination with oxygen and some metallic element. It is in fact a most valuable ingredient in the soil, for plants cannot get on without it. They require it especially for their fruits and seeds, in which the element, still in a state of combination, is stored up. From this store the animal world absorb large quantities of phosphorus, which is to them no less essential than it is to plants. There seems to be always a certain amount of phosphorus in the substance of which the brain is composed. But it is in the bones that the largest amount of phosphorus is found. In these bones it exists, in combination with oxygen and calcium, as phosphate of lime, $Ca_3 (PO_4)_2$. If bones be burnt the animal matter in the bones is got rid of, and the bone-ash that remains consists almost entirely of this phosphate of lime. The bone-ash is treated with dilute sulphuric acid, and the clear liquid is separated from the white powdery substance which is produced. The clear liquid is then boiled down to a syrup, and heated with charcoal in an iron retort. Phosphorus distils over, and must be condensed under water. The phosphorus so obtained is a white or light yellow solid, tolerably soft, and not unlike wax in appearance.

Let us take a small fragment of phosphorus, and tie it to the end of a piece of cotton. As it hangs in the air the phosphorus gives off a heavy white smoke, which seems to fall downwards from the little piece of yellow substance. If this be done in the dark the phosphorus is seen to give out a peculiar pale light. Hence the name phosphorus, or light-bearer. The phosphorus is all the while slowly uniting with the oxygen of the air; and the heavy white smoke is an oxide of phosphorus, called phosphorus trioxide, containing two atoms of phosphorus combined with three atoms of oxygen, $P_2 O_3$. If we hang the phosphorus over the mouth of a test tube, at the bottom of which there is some water coloured with blue litmus, the trioxide will dissolve in the water, and the litmus will be reddened. This shows that an acid is produced. It is called phosphorous acid, $H_3 PO_3$, and is formed thus:

$$P_2 O_3 + 3 OH_2 = 2 H_3 PO_3.$$

One molecule of phosphorus trioxide dissolves in and combines with three of water to produce two molecules of phosphorous acid.

Phosphorus very readily catches fire. Great care must therefore be taken in dealing with this element. It should always be cut under water. When it burns in the air a thick white smoke is produced. This is also an oxide of phosphorus, but not the same oxide. In it two atoms of phosphorus are combined with five atoms of oxygen. It is called phosphorus pentoxide, $P_2 O_5$. From this we may see that the same element may unite with different proportions of oxygen to form different compounds. This oxide of phosphorus also dissolves in water to form an acid. But this acid is of course different to that formed when the other oxide dissolves in water: it is called phosphoric acid, $H_3 PO_4$, and is formed thus:

$$P_2 O_5 + 3 OH_2 = 2 H_3 PO_4.$$

One molecule of phosphorus pentoxide unites with three of water to form two of phosphoric acid.

Ordinary yellow phosphorus may, like sulphur, be dissolved in carbon disulphide; but if we let the carbon disulphide evaporate, we shall not so readily obtain crystals of phosphorus as we obtained crystals of sulphur. The globules of phosphorus we get are light yellow in colour; but if we expose them to bright sunshine, they turn reddish, and then do not altogether dissolve in carbon disulphide. Let us place a little pellet of phosphorus at the bottom of this dry test tube, and then cork the test tube. On gently warming the bottom of the tube (*over a plate full of water for fear the tube should crack*) the phosphorus catches fire. It soon combines with all the oxygen in the test tube, and then goes out. If now we warm the phosphorus that remains, some of it passes off as vapour, and condenses in yellow drops in the cooler part of the tube, but some of it turns red. On letting the tube cool, and pouring in some carbon disulphide, we find that the yellow drops in the tube dissolve; but the phosphorus that has turned red does not dissolve. We can get some of this red phosphorus in another way. I place a small pellet of the ordinary yellow phosphorus in the middle of a plate, and then place on it a minute fragment of solid iodine. After a moment the two elements begin to unite, and the phosphorus bursts into flame. I immediately quench the flame by pouring water on the phosphorus. We now see that the phosphorus that remains is red. On treating it with carbon disulphide the red phosphorus does not dissolve. On examining this red phosphorus we find that it is not very inflammable, and does not give off white smoke in the air. On heating it in a closed test tube (from which the oxygen has been burnt out), we find that it can be distilled off, and that it condenses in minute drops of yellow phosphorus.

We may make one more experiment. Let us take a *very* minute fragment of red phosphorus, and a very minute fragment of potassium chlorate. I place them on this piece of iron, and give them a gentle blow with a hammer. There is a slight explosion, and a little cloud of smoke is formed. The phosphorus steals oxygen from the potassium chlorate, and unites with it to form phosphorus pentoxide. In our safety-matches the tips of the matches are composed of a substance which contains, besides inflammable matter, potassium chlorate. On the box there is spread a preparation which contains red phosphorus. When we strike a match, in the ordinary way, we bring the potassium chlorate into forcible contact with the red phosphorus; a slight explosion follows, and the inflammable matter on the tip of the match is thus ignited.

XXXV. OZONE AND NITRIC ACID.

We have seen that phosphorus, sulphur, and carbon are capable of existing in more than one state or condition. Phosphorus, for instance, exists either as ordinary yellow phosphorus or as red phosphorus. Sulphur may be obtained either in the ordinary brittle state, or in the elastic india-rubber-like condition. Carbon may be in the form of diamond, of graphite, or of charcoal. Each of these elements is therefore said to exist in several **allotropic** states. We cannot say for certain how these allotropic conditions are brought about. It is possible, however, that the different allotropic states correspond to different atomic groupings in the molecule.

There is a very remarkable allotropic modification of oxygen, which is known as **ozone**. We shall only be able to make it in very small quantities, mixed with much air or ordinary oxygen. It will be well, therefore, to have some means of testing for the presence of even a small quantity of ozone. This we can easily manage. I dissolve a little starch in warm water, and then add a few drops of a solution of a salt called potassium iodide, and stir the two together until they are well mixed. A piece of paper moistened with this solution of starch and potassium iodide will turn blue, or deep violet, if any ozone comes in contact with it. I now put two pieces of blotting paper so moistened into a glass cylinder, to the sides of which they adhere. At the bottom of the cylinder there are a few drops of warm water to keep the air in the cylinder moist. I now hang in the cylinder a piece of phosphorus which I have tied to the end of a thread of cotton. The phosphorus slowly unites with some of the oxygen, phosphorus trioxide being formed as a heavy white smoke. But some of the oxygen is converted into ozone, as is shown by the blue coloration imparted to our test papers. We can make a little ozone by pouring concentrated sulphuric acid upon some of these delicate violet crystals. They are crystals of a salt called potassium permanganate, which some of us know in the form of the beautiful and useful solution known as Condy's Fluid. By the action of the sulphuric acid oxygen is given off. But the oxygen is mixed with ozone, as is shown by the blue coloration of our test paper. Note that the violet crystals have dissolved in the sulphuric acid; but that the solution is not violet, but deep green. It is not simply a solution of the permanganate; it is a new substance, called manganese heptoxide. The ozone has a curious smell. It is half as heavy again as oxygen, its density being 24; and it oxidizes very powerfully, rapidly tarnishing, for example, the surface of mercury. It is produced when a series of electric sparks are passed through oxygen. Now there is very little—if any—doubt that the difference between ozone and oxygen is this: that in ozone there are three atoms in the molecule, while in oxygen there are only two. That is why I say it is possible that the difference in the allotropic conditions of phosphorus, sulphur, and carbon, are due to variations in the number of atoms in the molecule.

Ozone is present in small quantities in the atmosphere. If we hang out our ozone test papers in the air, we shall probably—unless we live in the neighbourhood of a smoky town—be able to detect its presence in the air. Unfortunately, however, there are other substances which have the same effect as ozone on our test papers. Some of these are the higher oxides of nitrogen, which are produced in the air by the action of lightning: so that some of the effect upon our test paper may be due to this cause.

If a series of electric sparks be passed through **dry** air, contained in a glass globe, brown fumes of these higher oxides of nitrogen are formed; and these brown fumes rapidly turn our starch and potassium iodide papers blue. In this experiment we imitate on a small scale what the lightning does on a large scale. If the glass globe contain **damp** air, blue litmus papers are turned red. This shows that an acid is produced. The acid is called **nitric acid**.

Nitric acid is a compound containing hydrogen, nitrogen, and oxygen, in the proportions of one part by weight of hydrogen, fourteen parts by weight of nitrogen, and forty-eight parts by weight of oxygen. The shorthand formula for it therefore is HNO_3. It may be made, as we have just seen, by passing electric sparks through damp air; also by passing electric sparks through a mixture of its three constituent gases. But as we very probably have no opportunities of performing experiments involving the passage of a series of electric sparks, it will be better for us to prepare the acid by a somewhat roundabout process. Into this curved vessel, called a retort, we must put some nitrate of potassium, or nitre, as it is generally called. On to this we must then pour enough sulphuric acid just to cover it. On heating these substances gently in the retort, drops of liquid are seen to collect in the straight part of the vessel. The end of the retort must be placed in a receiving flask, kept cool by being placed in an evaporating basin containing water, and being covered with a piece of damp blotting paper. Into this receiver the drops run. Soon we shall have collected a small quantity of the nitric acid, which appears to be a yellow liquid. The yellow colour is, however, due to an impurity, for pure nitric acid is colourless. On allowing the retort to cool, we find a white solid substance left. It is called sulphate of potassium. We may express the reaction in an equation thus:—

$$2\,KNO_3 + H_2\,SO_4 = 2\,HNO_3 + K_2\,SO_4.$$

Two molecules of nitre, with one molecule of sulphuric acid, yield two molecules of nitric acid and one of potassium sulphate.

Let us notice that the potassium (K) and the hydrogen (H) just change places.

We can readily see that the nitric acid so produced is an acid, for it turns blue litmus red. It is a very strong acid. If we pour some of it on to some metallic copper the metal is at once attacked. Dark brown fumes are formed; and these fumes will turn our starch and potassium iodide paper deep blue—they are fumes of one of the higher oxides of nitrogen. The copper disappears, and we have a blue solution containing copper nitrate. Lead and silver both dissolve in nitric acid, nitrate of lead or nitrate of silver being formed. Nitric acid will, indeed, dissolve all the commoner metals except gold and platinum. If a drop of the acid falls on your skin it will turn brown; if it falls on your black coat sleeve it will leave a red spot. Nitric acid, in fact, very readily eats into and corrodes animal or vegetable tissues.

XXXVI. SOME COMPOUNDS OF NITROGEN.

We have already seen that nitric acid is a compound containing nitrogen, hydrogen, and oxygen. We have also seen that if we pour nitric acid on metallic copper, brown fumes are produced. Let us now do this a little more carefully. Let us place some copper turnings—thin strips of metallic copper—in a flask like that we used for making hydrogen, and cover them with water. We may now add a few drops of concentrated nitric acid, and close the flask with a cork, through which there passes a glass tube, by which the gas produced may be led off for collection.

Brown fumes at once rise into the flask. But gradually the brownish-red colour in the flask fades away. And if we collect some of the gas that is being rapidly given off by the action of the nitric acid on the copper, allowing it to bubble up into a glass jar over water, we shall find that it is quite colourless. The reason of these colour-changes is seen the moment we admit some air into the glass jar which contains the colourless gas. For directly the air gains admittance, the brownish-red fumes again make their appearance. The gas produced by the action of nitric acid on metallic copper is, in fact, an oxide of nitrogen, having a composition represented by the formula NO. This gas is colourless; but the moment it comes in contact with atmospheric oxygen, it takes up an additional atom of oxygen, and becomes a higher oxide, represented by the formula NO_2. This oxide is ruddy-brown in colour. The former or lower oxide is called **nitric oxide**; the latter or higher oxide is called **nitric peroxide**. We thus see that just as phosphorus unites with different proportions of oxygen to form different oxides, so too does nitrogen. But in the case of nitrogen there are five different oxides. It will be sufficient for us to confine our attention to the nitric oxide and the nitric peroxide.

If we fill a glass jar three-quarters full of the colourless nitric oxide, and then blow in enough air to fill the rest of the jar, red fumes of nitric peroxide will be formed. But these red fumes quickly disappear, and the water rises in the jar. The ruddy nitric peroxide, in fact, dissolves very readily in water, whereas the colourless nitric oxide does not so dissolve. We may the experiment in the following manner. We may fill a glass jar half full of air, and then let nitric oxide slowly bubble up. As the bubbles of gas enter the jar, red-brown fumes of nitric peroxide are formed, but dissolve at once in the water. Thus the jar not only gets no fuller, but even becomes less full of gas. For the nitric oxide steals the oxygen from the air in the jar, and carries this oxygen into solution. If the experiment be stopped at the right moment nothing but nitrogen will remain in the jar.

Here are the equations for the preparation of nitric oxide, and its conversion into nitric peroxide:—

$$(1)\ 8\,HNO_3 + 3\,Cu = 3\,Cu\,(NO_3)_2 + 2\,NO + 4\,OH_2.$$

Eight molecules of nitric acid, with three atoms of copper, give three molecules of copper nitrate, two of nitric oxide, and four of water.

$$(2)\qquad 2\,NO + O_2 = 2\,NO_2.$$

Two molecules of nitric oxide combine with one of oxygen to form two of nitric peroxide.

We may now leave the compounds of nitrogen and oxygen to consider a compound of nitrogen and hydrogen. This compound is called **ammonia**. It consists of one atom of nitrogen combined with three atoms of hydrogen, and is therefore represented by the formula NH_3. It, too, like nitric acids, exists in small quantities in the atmosphere. In fact, it is generally found in the atmosphere in combination with nitric acid.

The two gases, nitrogen and hydrogen, cannot readily be made to unite directly, so that we are forced to prepare ammonia by a somewhat roundabout process. One way of preparing it is to heat the horns, hoofs, or hairs of animals with quicklime. These animal substances contain nitrogen and hydrogen, which are liberated in the form of ammonia by the action of the quicklime. When the gas thus prepared is dissolved in water, it goes by the characteristic name, **spirit of hartshorn.** Ammonia gas very readily unites with hydrochloric acid gas. To show this, pour a few drops of hartshorn into a tall glass jar and shake it up. The jar will be filled with the pungent-smelling ammonia gas. Cover the jar with a round plate of glass. Into a second jar pour a few drops of hydrochloric acid solution, and shake it up. The jar will be filled with the fumes of hydrochloric acid: Cover this jar with a plate of glass, invert it, and place it on the top of the other jar, mouth to mouth, and then remove the glass plates. A thick white smoke forms at once within the jars. The white smoke is produced by the union of the hydrochloric acid gas with the ammonia gas. A new compound called ammonium chloride ($NH_4\,Cl$) is formed. The white fumes of ammonium chloride are deposited on the sides of the glass vessels, and may be collected as a white powdery substance. If this white powdery substance be heated with quicklime ($Ca\,O$), the ammonia gas may be set free. We may represent the reaction in our chemical shorthand thus:

$$Ca\,O + 2\,NH_4\,Cl = Ca\,Cl_2 + OH_2 + 2\,NH_3.$$

One molecule of calcium oxide, with two of ammonium chloride, give one of calcium chloride, one of water, and two of ammonia gas.

Ammonia gas is very soluble in water. The gas cannot, therefore, be collected above water, as hydrogen or nitric oxide can. To show how soluble it is, we may fill a test tube with the gas. To do this, I place some of this strong solution of the gas in a test tube fitted with a cork, through which a short, straight glass tube passes. Over this tube I place an inverted test tube. I then heat the ammonia solution. When it is heated it can no longer hold so much of the gas in solution. The gas, therefore, rises rapidly into the upper test tube. When it has had quite time to become full of the gas, I may cover the mouth of the test tube with the thumb, and hold it beneath the surface of some water in a basin. On removing the thumb the water dissolves the gas, and rushes up into the test tube, completely—or almost completely—filling it with water. One thing we may especially notice about the solution of ammonia in water. It turns reddened litmus blue. Acids turn blue litmus red, but ammonia turns reddened litmus blue. Those substances which thus turn reddened litmus blue are called **alkalies** (see Lesson XIX.).

The densities of the three gases we have considered are: nitric oxide 15, nitric peroxide 23, and ammonia 8½. Let us here notice that in these cases (and in all others the same thing is found) the **density is just half the molecular weight**, or rather the number that expresses the density is just half the number that expresses the molecular weight. Let us show this in a table.

Name.	Formula of Mol.	Mol. wt.	Density.
Hydrogen	H_2	... 2	... 1
Oxygen	O_2	... 32	... 16
Ozone	O_3	... 48	... 24
Nitrogen	N_2	... 28	... 14
Ammonia	NH_3	... 17	... 8½
Nitric oxide	NO	... 30	... 15
Nitric peroxide	NO_2	... 46	... 23

Name.	Formula of Mol.	Mol. wt.	Density.
Water vapour	OH_2	18	9
Carbonic acid gas	CO_2	44	22
Sulphur dioxide	SO_2	64	32
Sulphuretted hydrogen	SH_2	34	17

This relation between the density and the molecular weight is partly due to the fact that hydrogen is the standard both for densities and for atomic weights. But the atomic weight of hydrogen being 1, the molecular weight is 2, for there are two atoms in the molecule. And since the density of hydrogen is 1, and the molecular weight 2, the latter is obviously twice the former. And if we take any other gas, such as ammonia, the molecular weight is 17 as compared with hydrogen 2, while the density is 8½ as compared with hydrogen 1. The ratio in each case is the same. There is, however, one more thing to notice. This relation would not hold good were it not for the fact that **the same volume of any gas always contains the same number of molecules**, so long as the temperature and pressure remain unchanged. It does not matter whether the gas be oxygen, hydrogen, ammonia, dry air, damp air, or a mixture of any or all of these, the number of molecules is exactly the same in any given space, say a cubic inch, so long as the temperature and pressure remains the same. Now in molecular weights we compare **single molecules**; and in densities we compare large **collections of molecules**, say in a cubic inch. But since the number of molecules in the cubic inch is always the same, the ratio for single molecules, say of hydrogen and ammonia, is the same as that for the equal numbers contained in a cubic inch.

XXXVII. THE CHLORINE GROUP.

We have seen that when two elements unite to form a chemical compound, they undergo a complete change of physical properties. We need not feel surprised, therefore, to learn that common table salt is composed of two elements, one of them a metal, and the other a poisonous gas. The metal is sodium; the poisonous gas is called **chlorine**. And just as the compound of a metal with oxygen is called an oxide of the metal, and the compound of a metal with sulphur is called a sulphide of the metal, so is the compound of a metal with chlorine called a chloride of the metal. Common salt is therefore called by chemists chloride of sodium. To separate the chlorine in this salt from the sodium with which it is combined, we must make use of sulphuric acid—an acid of which we shall learn something more in the next lesson; and we must mix some manganese dioxide with the sodium chloride which we treat with this acid. I have here such a mixture of sodium chloride and manganese dioxide; I put a little at the bottom of a test tube, and pour some concentrated sulphuric acid upon it. A yellowish green gas with an unpleasant smell is at once given off. It is a heavy gas, having a density of 35.5. This gas is the chlorine. The water over which we collect it must be warm, for chlorine dissolves in cold water. Or we may collect it by displacement, that is, we may let it flow into a glass jar covered with a paper lid, through which the glass tube from the test tube passes. If we hang in this tube a bunch of fresh flowers which have been dipped in water, we shall see that the colour quickly fades, until the flowers are white or yellowish white. Chlorine bleaches any animal or vegetable colour. To enable it to bleach, however, water must be present. The chlorine steals the hydrogen from the water, forming a substance called hydrochloric acid, and the oxygen thus set free attacks and destroys the colouring matter. Chlorine is also a valuable disinfectant, from its power of setting free oxygen from water when any noxious gases are present in the air.

If instead of taking the mixture of common salt and manganese dioxide we treat common salt alone with sulphuric acid, we shall not obtain chlorine, but a compound of chlorine with hydrogen, called hydrochloric acid gas. This gas is much lighter than chlorine, having a density of 18¼. It forms white fumes as it issues from the mouth of the test tube, especially if the air be damp. The invisible acid gas unites, in fact, with the invisible water-vapour in the air, and minute drops of a liquid solution of the gas in water are formed, and become visible as a white steam-like cloud. The pungent acid gas very readily dissolves in water, giving rise to a solution which used to be called "spirits of salt," but is now called, from the gas it holds in solution, hydrochloric acid. From this solution it is not difficult to obtain the chlorine. We may either pour it on manganese dioxide, or on potassium chlorate. In either case oxygen combines with the hydrogen of the hydrochloric acid, water being formed, and chlorine set free. Let us use potassium chlorate. I place a few crystals of this salt at the bottom of a test tube, and pour some hydrochloric acid upon them. Chlorine gas is rapidly given off.

This is perhaps the cleanest way of making the gas. The changes which take place may be seen by the following equation:

$$KClO_3 + 6\ HCl = 3\ OH_2 + 3\ Cl_2 + KCl.$$

One molecule of potassium chlorate, with six of hydrochloric acid, give three molecules of water, three of chlorine, and one of potassium chloride.

There are two elements which very closely resemble chlorine in some of their properties; they are called **bromine** and **iodine**. Their compounds with metals are called bromides and iodides. I have here three test tubes: each has a black substance at the bottom. I pour a little sulphuric acid upon each. In one a yellowish-green gas is given off. The gas is chlorine. In another a brownish-red vapour fills the lower part of the tube, and condenses in brown drops of a heavy liquid on the cooler parts of the test tube. This brown vapour which condenses to a liquid is the bromine. In the third a beautiful violet vapour fills the lower part of the tube, and condenses in the upper and cooler part of the tube into black, shining, scale-like particles. This violet vapour condensing to black solid particles is iodine. Each of the tubes contained manganese dioxide, mixed, in the first with potassium chloride, in the second with potassium bromide, in the third with potassium iodide. We may set down the three equations one under the other:—

$$(a)\ 2\ KCl + Mn\ O_2 + 2\ H_2\ SO_4 = K_2\ SO_4 + Mn\ SO_4 + 2\ OH_2 + Cl_2.$$

$$(b)\ 2\ KBr. + Mn\ O_2 + 2\ H_2\ SO_4 = K_2\ SO_4 + Mn\ SO_4 + 2\ OH_2 + Br_2.$$

$$(c)\ 2\ KI + Mn\ O_2 + 2\ H_2\ SO_4 = K_2\ SO_4 + Mn.SO_4 + 2\ OH_2 + I_2.$$

> Two molecules of (*a*) potassium chloride (*b*), potassium bromide, or (*c*) potassium iodide, with one molecule of manganese dioxide, and two of sulphuric acid, yield one of potassium sulphate, one of manganese sulphate, two of water, and one of (*a*) chlorine, (*b*) bromine, or (*c*) iodine.

I have here two test tubes, in one of which I have dissolved a minute crystal of potassium bromide, while in the other I have dissolved a minute crystal of potassium iodide. In a third test tube I make a little chlorine by pouring hydrochloric acid upon potassium chlorate. Chlorine is such a heavy gas, being nearly 2½ times as heavy as air, that we can pour it out of this test tube into those containing the solutions of the bromide and iodide. The upper part of the solution containing the bromide soon turns slightly yellow, while that containing the iodide turns yellowish-brown. What takes place is this: the chlorine seizes hold of the potassium, and unites with it in each case, leaving the bromine or the iodine, as the case may be, free and uncombined. I now pour into each test tube a few drops of carbon disulphide. The heavy drops sink to the bottom. But if I shake them up and down in the solution, they dissolve some of the bromine or iodine. The bromine colours the carbon disulphide reddish-brown; the iodine colours it a beautiful violet.

Both bromine and iodine form, with hydrogen, acids which resemble hydrochloric acid in several respects, and among others in the fact that they fume in damp air, that is, give rise to a cloud of minute droplets by uniting with the moisture invisibly contained in the air. They are called **hydrobromic acid** and **hydriodic acid** respectively.

There is also a fourth gas, which seems to belong to the same group of acid gases. It is called **hydrofluoric acid.** Written in our chemical shorthand, the four gases are: HCl, HBr, HI, and HFl. Hydrofluoric acid is formed by treating calcium fluoride (or fluor-spar, as the mineral is called) with concentrated sulphuric acid. The most curious property of this gas is its power of etching glass. To show this we must cover the piece of glass to be etched with wax. This is easily done by warming the glass, and smearing it over with fine white wax. I now take a fine-pointed instrument (the point of a file will do), and write the letters HFl on the waxed surface. The fine point removes some of the wax. Then place some pounded fluor-spar in this shallow leaden vessel, and pour upon it concentrated sulphuric acid. On warming the vessel, dense white fumes are formed. These fumes are produced by the union of the hydrofluoric acid gas with the invisible moisture of the air. In these fumes I suspend the waxed glass by means of a piece of string. The powerful acid attacks the glass where the point of the file removed the wax. Elsewhere the wax preserves the glass from the action of the acid. After a while we may remove the glass from the action of the acid fumes. On warming it the wax melts, and may be rubbed off with a cloth. The letters HFl are found to be etched on the surface of the glass.

In the manufacture of glass a good deal of sand is used. This sand is an oxide of the element silicon, SiO_2. The same substance is found in quartz, rock crystal, chalcedony, flint, agate, and other minerals. It is this silica, or oxide of silicon, in the glass that is attacked by the acid fumes, a new compound called fluoride of silicon being formed.

The elements, chlorine, bromine, iodine, and fluorine, seem to form a distinct group. Chlorine, bromine, and iodine may, as we have seen, be obtained in the elementary condition; but fluorine has not yet been isolated.

Note. As chlorine is poisonous, and hydrofluoric acid is irritating to the throat and lungs, the experiments above described should be done in the open air, or at an open window.

XXXVIII. SULPHURIC ACID.

We have seen that when sulphur burns in the air an oxide of sulphur (sulphur dioxide) is formed; and we have seen that this oxide of sulphur dissolves in water to form an acid called sulphurous acid. We have now to learn that sulphur forms another oxide, called sulphur trioxide, which combines with water to form another acid—namely, sulphuric acid. We have often had occasion to use this acid. Let us now see how it may be prepared. It is easy enough to make sulphur dioxide: all we have to do is to burn some sulphur in the air. But it is by no means so easy to obtain sulphur trioxide. The sulphur dioxide will not take up and combine with an additional atom of oxygen from the air. We can, indeed, cause sulphur dioxide and oxygen to unite, by passing them through a tube in which some metallic platinum, in a fine state of division, is heated to redness. But such spongy platinum, as it is called, is too expensive a substance to be used for the production of sulphuric acid upon a large scale, and for commercial purposes. It is necessary, therefore, to search for some chemical substance which will give up oxygen to the sulphur dioxide. Such a substance is nitric peroxide. Sulphuric acid is manufactured on a large scale in the following manner: Sulphur dioxide, formed either by burning crude sulphur, or by roasting iron pyrites, a mineral rich in sulphur, is led into a large leaden chamber. Into the same chamber steam is introduced, and a supply of atmospheric air is admitted; and into the same chamber nitric peroxide is allowed to pass. This gas is produced by allowing some concentrated sulphuric acid to act upon nitre, or nitrate of potassium. We have already seen that when these two substances are **gently** heated in a retort nitric acid is produced; but we have also seen that the nitric acid is apt to be coloured yellow by an impurity. That impurity is nitric peroxide. If, as in the case we are now considering, the acid and the nitre be more strongly heated together, nitric fumes are produced, of which nitric peroxide may be taken as the most important. No sooner does this nitric peroxide (to which, for the sake of simplicity, we will confine our attention) come in contact with sulphur dioxide, than it gives up an atom of oxygen to the sulphur dioxide, thus converting it into sulphur trioxide, while it is itself converted into nitric oxide. But, as we have already seen, nitric oxide, when it comes into contact with atmospheric air, unites at once with oxygen to form nitric peroxide. This reconstituted nitric peroxide is then ready to convert more sulphur dioxide into sulphur trioxide, being again converted into nitric oxide, which again takes up more oxygen from the air admitted into the leaden chamber, and thus again becomes nitric peroxide, which is in a position to oxidize more sulphur dioxide. We may express these changes in our chemical shorthand thus:—

$$(1)\ SO_2 + NO_2 = SO_3 + NO.$$
$$(2)\ 2\ NO + O_2 = 2\ NO_2.$$

The nitric peroxide (NO_2) thus acts as a carrier of oxygen from the air to the sulphur dioxide. It gives up oxygen to the sulphur dioxide, and then at once takes up oxygen from the air, and thus constantly hands on oxygen from the air to the sulphur dioxide. In the mean while the sulphur trioxide, thus produced, unites with the steam introduced into the leaden chamber, and with it forms sulphuric acid. For one molecule of sulphur trioxide (SO_3) combines with one of water (OH_2) to form one of sulphuric acid (H_2SO_4). If there is an insufficient supply of steam, white crystals, called "**crystals of the leaden chamber,**" are formed. These, however, are converted by water into sulphuric acid, ruddy fumes of an oxide of nitrogen being given off.

Let us now take the bottle we used for showing the formation of sulphur when sulphur dioxide and sulphuretted hydrogen come into contact. Through the cork there pass three tubes. Let us connect one with a flask in which nitric oxide is formed by the action of copper on nitric acid. By occasionally admitting air into this flask the nitric oxide is converted into nitric peroxide, the ruddy fumes of which pass into the wide-necked bottle. Another tube may be connected with a test tube, in which sulphur dioxide is produced by the action of sulphuric acid on sulphur. The third tube may be connected with a long glass tube, by which the waste fumes may pass out of the window into the open air. We shall find, when all our tubes are connected, that the inside of the bottle is rapidly frosted over with a white solid, which sometimes spreads in fern-like patterns. This white solid consists of crystals of the leaden chamber. When a sufficient quantity has been collected we may pour some water into the bottle. It will hiss as it comes in contact with and dissolves the crystals. Ruddy fumes of an oxide of nitrogen are given off, and the solution contains sulphuric acid.

Concentrated sulphuric acid is a thick, oily-looking liquid. It unites very vigorously with water. I pour a few drops of the strong acid into a test tube containing a few drops of water. The mixture becomes very hot. The affinity of sulphuric acid for water is so strong that it will extract water from substances in which the oxygen and hydrogen are chemically combined with other elements. Sugar, for example, contains oxygen and hydrogen in the right proportions to form water; but these elements are combined with a third element—namely, carbon. I pour into the bottom of this small beaker some strong syrup made from loaf sugar, and, having placed the beaker on a plate, pour into the syrup some concentrated sulphuric acid. The result is that the syrup begins to blacken and froth up over the sides of the beaker, until it becomes a mass of sugar charcoal. The sulphuric acid has robbed the sugar of its oxygen and hydrogen, so that only carbon is left. Sulphuric acid will also take away a molecule of water from alcohol. That which remains is, of course, no longer alcohol, but a gas which burns with a bright smoky flame, and which is indeed an important constituent of coal gas. If we dip a splinter of wood into strong sulphuric acid it will be charred. The acid robs the wood of water, and a black carbonaceous mass remains.

We have already seen how useful sulphuric acid is in enabling us to obtain other substances. By treating zinc with dilute sulphuric acid we obtained hydrogen. By treating copper with concentrated sulphuric acid we obtained sulphur dioxide. By treating nitre with sulphuric acid we produced nitric acid. With sodium chloride sulphuric acid gave rise to the production of hydrochloric acid: and with potassium permanganate crystals sulphuric acid gave rise to the formation of ozone. To these we may now add two more. If manganese dioxide is treated with sulphuric acid oxygen will be produced. We may carry out this experiment in the flask we used for hydrogen, and for carbonic acid gas. But if chlorate of potassium be treated with strong sulphuric acid a dark yellow substance called chlorine peroxide is produced. **Great** care must be taken in the preparation of this substance, as it is apt to explode with some violence. It oxidizes very powerfully. If a few small pellets of phosphorus, and a few crystals of potassium chlorate, be placed at the bottom of a large test tube, and then covered with water, we may pour sulphuric acid down the side on to the crystals, and we shall then see the phosphorus burn with bright flashes under water. It obtains the oxygen for its combustion from the chlorine peroxide. If we mix equal proportions of potassium chlorate and sugar (pounded up **separately** and then mixed carefully), and then touch the mixed mass with a drop of concentrated sulphuric acid, it will suddenly burst into flame. A little chlorine peroxide is formed; this gives up oxygen to a neighbouring particle of sugar, which bursts into flame; and then the flame spreads, the sugar burning brightly at the expense of the oxygen in the potassium chlorate. In much the same way a safety-match may be ignited by a touch of a glass rod moistened with concentrated sulphuric acid.

XXXIX. MODES OF CHEMICAL ACTION.

We will in this lesson consider some of the ways in which chemical action takes place. We may set first, as that mode of chemical action which is perhaps the simplest and most important, **direct chemical union.** Of this we have had many examples. To take one more: here is a piece of metallic magnesium, which has been prepared in the form of a ribbon-like strip. On holding one end in the lamp it burns with an extremely bright, somewhat blue light. A white smoke, consisting of magnesium oxide, is formed during the combustion. In this experiment the metal magnesium enters into direct chemical union, atom for atom, with the oxygen of the air, a compound (MgO) being formed. Sulphur also unites directly with oxygen, but in this case one atom of sulphur unites with two atoms of oxygen, the compound formed being, therefore, SO_2. To take one more case: phosphorus unites directly with the oxygen of the air; during the act of combustion two atoms of phosphorus unite with five atoms of oxygen, and the compound formed is P_2O_5. We may carry this experiment a step further. By burning the phosphorus on a plate under a large glass beaker (or a large glass ornament shade), we may collect a good deal of the phosphorus pentoxide. If we put some at the bottom of a dry test tube, and then pour in a little water, the oxide will unite chemically with some of the water; a slight hissing noise will accompany the union, and the test tube will become hot. Here then we have a case of two chemical compounds uniting together to form a more complex compound, this more complex compound being phosphoric acid. An equation will best show the mode of union:—

$$P_2 O_5 + 3 OH_2 = 2 H_3 PO_4.$$

The second mode of chemical action is **chemical decomposition.** Of this, too, we have had many examples. Let us repeat one of the experiments which illustrate this mode of action. I here heat some red mercury oxide in a glass tube. It turns black, and then begins to decompose. Oxygen is given off, and may be recognized by the fact that it rekindles a glowing match. Mercury is deposited as a ring of bright metallic globules in the cooler part of the tube. The compound mercury oxide here splits up into the elements mercury and oxygen. Sometimes, however, the compound splits up into an element and a simpler compound. Here for example is some red lead; in it three atoms of lead are united with four atoms of oxygen. I heat some at the bottom of this test tube. The red lead at first turns black; but the colour eventually changes from red to yellow; a simpler oxide of lead is produced, and oxygen is given off, which may be recognized by its power of re-igniting a glowing match. Sometimes the compound splits up into two simpler compounds. Of this we have had a good example in the action of heat upon carbonate of lime, which decomposes into oxide of lime and carbonic acid gas, the Ca CO_3 becoming CaO and CO_2. There are thus many phases of this mode of chemical action.

We now come to a different mode—namely, that in which one element or group of elements is substituted for another. Of this **substitution**, perhaps the best example is the action of the metal sodium on water. I throw this pellet of the light metal on to water in a small evaporating dish. The metal rolls about and gradually disappears, while hydrogen is given off. What takes place is this: an atom of the metal sodium is substituted for one atom of hydrogen in the water, the alkaline hydrate of sodium being formed. To take another example: here is a bottle in which sulphuretted hydrogen (H_2 S) was some days ago dissolved in water. I have purposely left the cork out, and stood it outſide in the daylight. The bottom of the bottle is thickly covered with a yellow powder. That yellow powder is sulphur, oxygen of the air having been substituted for the sulphur in the sulphuretted hydrogen, and water having been formed. One more illustration of this simple substitution may be given. I have here a solution containing lead. In it I hang, by a piece of cotton, a piece of metallic zinc. The result is that the zinc is substituted for the lead in the solution, and the lead is substituted for the zinc at the end of the piece of cotton. The lead collects there in a beautiful fern-like and branching crystalline form, so that it has received the name of the lead tree. A yet more beautiful case of metallic substitution is seen if a globule of mercury be placed in a solution of silver nitrate. Silver is substituted for some of the mercury, a bright, metallic, fern-like, crystalline mass being produced.

A little more complex, as a mode of chemical action, is **double or reciprocal substitution.** I have here some copper nitrate, produced by the action of nitric acid on copper; it forms a blue solution. Into it I pour some sodium hydrate, and a blue solid at once makes its appearance. If we filter off this precipitate, we shall find that the solution is now colourless. We have here a case of reciprocal substitution, or mutual interchange. The nature of the reaction is, perhaps, more readily seen by words than by symbols. It is this:

Copper nitrate and sodium hydrate give copper hydrate and sodium nitrate.

Take another example. I have here a solution of lead nitrate, and to it I add some sulphuric acid, or, as it is often called, hydrogen sulphate. A white precipitate is at once produced. The reaction is as follows:

Lead nitrate with hydrogen sulphate give hydrogen nitrate and lead sulphate.

Here again, therefore, we have reciprocal substitution. When we made nitric acid, or hydrogen nitrate, we had another instance of the same mode of chemical action. In that case the reaction was:

Potassium nitrate and hydrogen sulphate give hydrogen nitrate and potassium sulphate.

Here again there is double substitution.

We may now notice that all acids contain hydrogen. The only apparent exception is carbonic acid gas; and this is not a true acid until it combines chemically with water, in which case it too becomes a true acid, containing hydrogen. One of the commonest kind of chemical reactions is that in which the hydrogen of an acid is replaced by a metal, the metal being substituted for the hydrogen. When this takes place, in any case, a **salt** is formed. Thus the lead nitrate and the lead sulphate, the copper nitrate and the sodium nitrate, are examples of salts obtained from the acids, hydrogen nitrate and hydrogen sulphate, by the substitution of the metal, lead, copper, or sodium, for the element hydrogen. Here is some dilute sulphuric acid, and into it I drop some pieces of metallic zinc. Hydrogen is at once given off, and zinc is substituted for it, the salt zinc sulphate being formed.

One more thing may be noticed before closing this lesson. It is this: that the atoms of the various elements have various **substitution values.** One atom of sodium, for example, is always substituted for one atom of hydrogen, never one of sodium for two of hydrogen. On the other hand, one atom of zinc is always substituted for two atoms of hydrogen, never one of zinc for one of hydrogen. Thus the symbol of sodium sulphate is $Na_2\,SO_4$, while that of zinc sulphate is $Zn\,SO_4$, both being derived from hydrogen sulphate, $H_2\,SO_4$. Copper and lead are like zinc in this respect, one atom of each of these metals being substituted for two atoms of hydrogen; but potassium is like sodium, and is substituted for hydrogen, atom for atom. Bearing this in mind, we may now, in conclusion, set down the shorthand equations for those reactions which we have above expressed in words.

1. For the action of sodium hydrate on copper nitrate:

$$2\,Na\,HO + Cu\,(NO_3)_2 = Cu\,H_2O_2 + 2\,Na\,NO_3.$$

2. For the action of hydrogen sulphate on lead nitrate:

$$H_2\,SO_4 + Pb\,(NO_3)_2 = Pb\,SO_4 + 2\,H\,NO_3.$$

3. For the action of hydrogen sulphate on potassium nitrate:

$$H_2\,SO_4 + 2\,K\,NO_3 = K_2\,SO_4 + 2\,H\,NO_3.$$

XL. MORE ABOUT MOLECULES.

MUCH has been said in previous lessons on molecules, and a molecule has been defined as the smallest particle of any substance which retains all the essential properties of that substance. Let us take a drop of water, for example. That drop is made up of an immense number of minute molecules. In water, as in any **liquid**, the molecules are in constant rolling motion over each other. It is the energy of the heat that becomes latent in water which enables these molecules to continue their rolling motion. Take away that energy, and the rolling motion ceases; the molecules are locked together, as it were, and the liquid water becomes **solid** ice. If, however, the water be gradually heated to a temperature of 100° C., the rolling motion of the molecules becomes more and more violent, until at 100° the molecules begin to fly apart from each other, and the water becomes vapour of water. In that vapour, as in any **gaseous** substance, the molecules are dashing about with extreme rapidity, and constantly knocking against each other, and against the sides of any vessel in which they are enclosed. This constant bombardment of the sides of the vessel gives rise to the **pressure** exerted by the gas. We have seen that the same volume of any gas under the same conditions of pressure and temperature contains the same number of molecules. We have seen, too, that if the volume of a gas be forcibly diminished, the pressure exerted by the gas is increased. This is because a greater number of molecules are forced into the space that was occupied before by a less number of molecules, and the bombardment of a greater number of molecules, moving at the same rate, gives an increased pressure. But if the temperature of a gas be increased, the effect on the molecules is to increase the rapidity of their motion. If, then, the gas in a certain enclosed space be heated, the molecules move more rapidly, and the more vigorous bombardment of the molecules, dashing more violently against the sides of the vessel, produces an increase of pressure.

Simple illustrations of the fact, that both in liquids and in gases there is a constant motion among the molecules, are afforded by the following well-known facts of observation. If a drop or two of coloured liquid be added to a tumbler of water the colour is rapidly diffused throughout the whole mass of water in the tumbler. If the stopper of a bottle containing some strong scent be removed, the scented particles become diffused throughout the air of the room.

We have seen that in the case of water a certain amount (79 thermal units) of heat becomes latent when the water passes from the solid to the liquid state, and again, that a further amount (537 thermal units) of heat becomes latent when it passes from the liquid into the gaseous condition (see Appendix p.

(145). This latent heat is insensible to the thermometer. We may say that the heat is rendered insensible in the one case, because it is fully occupied in keeping the molecules in rolling liquid motion, and that it is insensible in the other case because it is fully occupied in keeping up the free gaseous motion of the molecules. The energy is, in fact, so fully occupied in keeping the molecules in motion that it is, so to speak, too busy to take any notice of the thermometer.

But it may be asked, what proof have we that, in a vessel filled, for example, with water vapour at 100° C., there are a great number of separate molecules, and that the vessel is not continuously filled with water-vapour? How, in fact, do we know that there are **separate** molecules, with space between through which the molecules are in constant motion? In this way. If a large glass globe be kept at a temperature of 100° C., and be connected with a small vessel in which water is heated to 100° C., water molecules will rise and fill the glass globe. After a while the globe will be as full as it can hold of water molecules. No more will rise from the small vessel so long as the temperature remains at 100° C. But if now a glass vessel containing spirits of wine be connected with the globe, and gradually heated to 100° C., spirits of wine molecules will rise into the globe. In fact, just as many spirits of wine molecules pass into the globe as if there were no water molecules there at all. The only explanation of this is that the spirits of wine molecules get into the space **between the water molecules**. The pressure on the sides of the globe is, however, increased, for it is due to the double bombardment of the water molecules **and** the spirits of wine molecules. And if, when no more spirits of wine molecules will rise into the globe, a glass vessel containing another liquid, such as ether, be connected, and gradually heated to 100° C., as many ether molecules will rise into the globe as if there were no water molecules, and no spirits of wine molecules. They, too, can find plenty of room in the space between the molecules.

APPENDIX.

SOME arithmetical questions, based on the preceding lessons, are here appended. They (and more besides) should be worked out by all who have the requisite arithmetical knowledge.

1. Convert (*a*) 60° C., (*b*)—29° C., (*c*)—40° C., (*d*) 200° C., into F. (see L. XII.). Answers: (*a*) 140° F., (*b*)—20.2° F., (*c*)—40° F., (*d*) 392° F.

2. Convert (*a*) 50° F., (*b*)—20° F., (*c*)—4° F., (*d*) 14° F., into C. (L. XII.). Answers: (*a*) 10° C., (*b*)—28.8 C., (*c*)—20° C., (*d*)—10° C.

3. How many cubic inches of ice will be formed when 15 cubic inches of water are frozen? (NOTE. One cubic inch of water becomes $1\frac{1}{11}$ cubic inch of ice.) Ans.: $16\frac{4}{11}$ cubic inches.

4. If 60 cubic inches of ice are melted, how many cubic inches of water will be formed? Ans.: 55 cubic inches.

5. How many cubic feet of ice will be formed if 121 cubic feet of water be frozen? Ans.: 132 cubic feet.

6. How many cubic inches of water will result from the melting of a block of ice, 18 inches long, 8 inches broad, and 2 inches deep? Ans.: 264 cubic inches.

7. If 5 cubic inches of water be converted into vapour of water, how many cubic inches of this vapour will there be? (**Note.** One cubic inch of water becomes 1700 cubic inches of water-vapour.) Ans.: 8500 cubic inches.

8. If 34 cubic inches of water-vapour be condensed, how much water will be produced? Ans.: .02 cubic inch.

9. How much water must be converted into vapour in order that a space containing 425 cubic inches may be filled with the vapour? Ans.: .25 cubic inch.

Introductory note to problems 10—24.

The latent heat of water is 79 thermal units; that is to say, to melt one pound of ice as much heat is required as would raise 79 pounds of water 1° C. (a thermal unit being the amount required to raise one pound of water 1° C.). Suppose we mix one pound of water at 79° C. with one pound of ice at 0°. The water contains 79 thermal units, and this is just sufficient to melt the ice. The temperature of the mixture will therefore be 0°. Suppose we take two pounds of water at 39.5° and mix them with one pound of ice at 0°. Each pound of water at 39.5 contains 39.5 thermal units: the two pounds of water therefore contain twice 39.5, or 79 thermal units, and this is just sufficient to melt the ice. The temperature of the mixture will be 0°. Once more, suppose we have one pound of water at 81° and mix it with one pound of ice at 0°. The pound of warm water contains 81 thermal units. Of these 79 go to melt the ice. There remain, in the pound of water which was at

81°, two thermal units. We have, therefore, this pound of water at 2° mixed with one pound of water at 0° resulting from the melting of the ice. There two pounds of water intermingle, and share the two thermal units between them. So that there remain two pounds of water at 1°.

The latent heat of water-vapour is 537 thermal units. Suppose we pass half a pound of steam into ice-cold water. Let us see what weight of water this half pound of steam would raise to the boiling point. One pound of steam gives out 537 thermal units; half a pound will therefore give out 268.5 thermal units; and this amount of heat would raise 2.685 pounds of water from 0° to 100°.

10. How many pounds of water at 15.8° C. will be required just to melt one pound of ice? Ans.: 5 pounds.

11. One pound of water at 85° C. is mixed with one pound of ice at 0° C. What will be the final temperature when all the ice is melted, and all the water has the same temperature? Ans.: 3° C.

12. How much water at 31.6° C. will be required to melt .3 lb. of ice? Ans.: .75 lb.

13. How much steam (vapour of water at 100°) will have to be condensed in 10.72 lbs. of ice-cold water to raise them to the boiling point? Ans.: 2 lbs.

14. How much heat must pass into 5 lbs. of water at 100° C. in order to convert them into steam? Ans.: 2685 thermal units.

15. A piece of a certain substance weighs 15 grains in the air, and 12 grains in water. Find its specific gravity. (L. xxx.) Ans.: 5.

16. A piece of silver weighs 42 grains in the air. What will it weigh in water if its specific gravity is 10.5? Ans.: 38 grains.

17. A piece of sulphur weighs 1.23 gram in the air, and .63 gram in water. What is its specific gravity? Ans.: 2.05.

18. What is the specific gravity of a substance the weight of which in air is to its weight in water as 7 is to 5? Ans.: 3.5.

19. A boy can just lift a mass of iron weighing 46 pounds. What weight of iron could he just lift from the bottom of a tank full of water? The specific gravity of iron is 7.8. Ans.: 52.76 lbs.

20. Copper has a specific gravity of 8.9. What percentage of its weight in air does it lose when it is immersed in water? Ans.: 11.2°/₀.

21. The specific gravity of powdered tin-ore is to the specific gravity of the rocky matter with which it is mixed as 6.3 to 3. What will be their relative weights in water? Ans.: 5.3 to 2.

23. A piece of sulphur weighs 8 grains in the air, 4 grains in water, and 3 grains in another liquid. What is the specific gravity of this other liquid? Ans.: 1.25.

24. A mineral crystal weighs 1.48 gram in the air, 1.23 in water, and 1.255 in alcohol. What is the specific gravity of the alcohol? Ans.: .9.

Introductory note on the remaining problems.

Connected with the symbol of each chemical element there is a number which represents the relative weight of the atom of that element as compared with that of the atom of hydrogen which is taken as the standard.

The atom of hydrogen is said to weigh one **microcrith** (m.c.). The atom of oxygen weighs 16 m.c., that of sulphur 32 m.c., that of zinc 65 m.c., and so on. Remembering this, let us consider the equation which represents the action of zinc on sulphuric acid and the consequent setting free of hydrogen. That equation is—

$$Zn + H_2SO_4 = ZnSO_4 + H_2.$$

Now by reference to the table on page 45, we can ascertain the weight of any of these atoms. The atom of zinc weighs 65 m.c. Next to the zinc, in the equation, comes a molecule of sulphuric acid. To find the weight of this molecule we must add together the weights of all the atoms which compose it; and we must do the same to obtain the weight of the molecule of zinc sulphate. That of sulphuric acid is—

$$(H_2 =)\ 2 + (S =)\ 32) + (O_4 = 64) = 98 \text{ m.c.}$$

That of zinc sulphate is—

$$(Zn =)\ 65 + (S =)\ 32 + (O_4 =)\ 64 = 161 \text{ m.c.}$$

The molecule of hydrogen, consisting as it does of two atoms, each of which weighs one m.c., weighs two m.c. Let us now write the equation again with the weights in m.c. underneath.

$$\begin{array}{ccccccc} Zn & + & H_2SO_4 & = & ZnSO_4 & + & H_2. \\ 65 & + & 98 & = & 161 & + & 2. \end{array}$$

Notice that the weights to the left, added together, are equal to the weights to the right, added together; 163 in each case. We see, then, that to obtain two microcriths of hydrogen we must take 65 m.c. of zinc. So, too, if we wish to obtain two **grains** of hydrogen, we must take 65 grains of zinc; and if we wish to obtain two pounds, or grams, or ounces of hydrogen, we must take 65 pounds, or grams, or ounces of zinc; and so on. We are now able by the use of "proportion" to answer a number of arithmetical questions. Suppose, for example, we wish to fill a small balloon with hydrogen, and we know that four grains of hydrogen will just fill the balloon. We wish therefore to know (*a*) how much zinc and (*b*) how much sulphuric acid we must use. Then we have these rule of three sums:

a. If to get two grains of hydrogen we require 65 grains of zinc (by the equation), how much zinc shall we require to get four grains of hydrogen? Stating in proportion thus:

2 : 65 : : 4 : answer,

we find the answer to be 130 grains.

b. If to get two grains of hydrogen we require 98 grains of sulphuric acid (by the equation) how much sulphuric acid shall we require to get four grains of hydrogen? Again, stating in proportion, thus:

2 : 98 : : 4 : answer,

we find the answer 196 grains.

And suppose we wish to know how much hydrogen we can get by dissolving 32½ grams of zinc in sulphuric acid, we have this rule of three sum:

If 65 grams of zinc give two grams of hydrogen (by the equation), how much hydrogen will 32½ grams of zinc give? Stating in proportion, thus:

65 : 2 : : 32.5 : answer,

we obtain one gram of hydrogen as the answer.

We will take one more case. Suppose we wish to obtain an answer to the following question: How much potassium chlorate must be heated with manganese dioxide to obtain 19.2 grams of oxygen gas? Now we know that the manganese dioxide does not change, so we may leave that out of our equation, which will then be as follows:

$$\begin{array}{lcll} 2\ KClO_3\ (\text{heated with } MnO_2) & = & 2\ KClO & +\ 3\ O_2. \\ 2 \times 122.5 & & 2 \times 74.5 & \ \ 3 \times 32. \\ \quad 245 & = & \quad 149 & +\ \ 96. \end{array}$$

From this we see that to obtain 96 grams of oxygen, we must take 245 grams of potassium chlorate. We may, therefore, state in proportion thus:

$$96 : 245 : : 19.2 : \text{answer},$$

by which we find the answer to be 49 grams.

The problems which follow are similar in character to those which have just been considered. The equation on which the problem is based is given in each case.

25. How much zinc must be dissolved in sulphuric acid in order that five grams of hydrogen gas may be produced?

$$Zn + H_2\ SO_4 = Zn\ \ SO_4\ + H_2.$$

Ans.: 162.5 grams.

26. How much hydrogen by weight will be produced by dissolving 3.75 grams of zinc in sulphuric acid? Ans.: .115 gram, or 115 milligrams.

27. If 13 grains of zinc be dissolved in sulphuric acid, how much zinc sulphate will be produced? Ans.: 32.2 grains.

28. How much oxygen can be obtained by heating 12.25 grains of potassium chlorate?

$$2\ KClO_3 = 2\ KCl + 3\ O_2.$$

Ans.: 4.8 grains.

29. How much potassium chlorate must be heated to obtain 1.6 decigram of oxygen? Ans.: 4.08 decigrams.

30. How much mercury oxide must be heated to obtain two grains of oxygen?

$$2\ Hg\ O = 2\ Hg + O_2.$$

Ans.: 27 grains.

31. How much copper must be dissolved in sulphuric acid to give eight grams of sulphur dioxide?

$$Cu + 2\ H_2\ SO_4 = Cu\ SO_4 + 2\ OH_2 + SO_2.$$

Ans.: 7.875 grams.

32. How much sulphur dioxide will be produced by dissolving 126 centigrams of copper in sulphuric acid? Ans.: 128 centigrams.

33. If 100 grams of copper are acted upon with 50 decagrams of sulphuric acid, how much (*a*) sulphuric acid will remain unchanged, and how much (*b*) coppe, sulphate and (*c*) sulphur dioxide will be produced? Ans.: (*a*) $188\frac{8}{9}$ grams, (*b*) 252.38 grams, (*c*) 101.59 (nearly) grams.

34. How much water will be required to convert eight grams of sulphur dioxide into sulphurous acid?

$$SO_2 + OH_2 = H_2\ SO_3.$$

Ans.: 2.25 grams.

35. How much phosphorus pentoxide will be produced by burning 6.2 grains of phosphorus in the air?

$$4\ P + O_2 = 2\ P_2\ O_5.$$

Ans.: 14.2 grains.

36. How much phosphorus trioxide must be dissolved in water to form 41 decigrams of phosphorus acid?

$$P_2\ O_3 + 3\ OH_2 = 2\ H_3\ PO_3.$$

Ans.: 27.5 decigrams.

37. How much phosphoric acid can be obtained by dissolving 71 grams of phosphorus pentoxide in water?

$$P_2\ O_5 + 3\ OH_2 = 2\ H_3\ PO_4.$$

Ans.: 98 grams.

38. How much sulphuric acid, and how much potassium nitrate, must be used in the preparation of seven grams of nitric acid?

$$2\ KNO_3 + H_2\ SO_4 = K_2\ SO_4 + 2\ HNO_3.$$

Ans.: 5.4 grams $H_2\ SO_4$
11.2 ,, KNO_3.

39. If 5.15 grains of potassium nitrate are treated with sulphuric acid, how much potassium sulphate will be formed? Ans.: 4.35 grains.

40. How much copper and nitric acid must be used to get 10 grains of nitric oxide?

$$3\ Cu + 8\ HNO_3 = 3\ Cu\ (NO_3)_2 + 2\ NO + 4\ OH_2.$$

Ans.: 31.5 Cu : 84 HNO_3.

41. If 1.5 gram of nitric oxide are converted by

the action of oxygen into nitric peroxide, how many grams of the latter gas will be formed?

$$2\ NO + O_2 = 2\ NO_2.$$

Ans.: 2.3 grams.

42. How much ammonium chloride must be heated with calcium oxide to obtain 20 grams of ammonia gas?

$$2\ NH_4\ Cl + Ca\ O = 2\ NH_3 + Ca\ Cl_2 + OH_2.$$

Ans.: 62.94 grams.

43. How much ammonia gas will be obtained by treating .2675 grain of ammonium chloride with quicklime? Ans.: .085 gram.

44. How much manganese dioxide must be used with sodium chloride and sulphuric acid to obtain 142 grams of chlorine?

$$2\ Na\,Cl + Mn\,O_2 + H_2\,SO_4 = Na_2\,SO_4 + Mn\,SO_4 + 2\,OH_2 + Cl_2.$$

Ans: 174 grams.

45. How much hydrochloric acid gas will be obtained by acting upon 468 milligrams of sodium chloride with sulphuric acid?

$$2\ Na\ Cl + H_2\ SO_4 = 2\ H\ Cl. + Na_2\ SO_4.$$

Ans.: 292 m.g.

46. How much chlorine can be obtained by acting upon 6.125 grains of potassium chlorate with hydrochloric acid?

$$KClO_3 + 6\ HCl = KCl + 3\ OH_2 + 3\ Cl_2.$$

Ans.: 10.65 grains.

47. How many centigrams of bromine will be produced by acting upon 20.6 grams of sodium bromide

with sulphuric acid, and how much manganese dioxide must be mixed with the sodium bromide?

$$2\ Na\,Br + Mn\,O_2 + H_2\,SO_4 = Na_2\,SO_4 + Mn\,SO_4 + 2\ OH_2 \times 2\ Br.$$

Ans.: 1600 c.g. Br 8.7 gr. MnO_2.

48. How many decagrams of iodine will be produced by acting upon 598 grams of sodium iodide with sulphuric acid, and how much manganese dioxide must be mixed with the sodium iodide?

$$2\ Na\,I + Mn\,O_2 + H_2\,SO_4 = Na_2\,SO_4 + Mn\,SO_4 + 2\ OH_2 + 2\ I.$$

Ans.: 50.6 decagrams I, 17.4 grams $Mn\,O_2$.

49. What will be the result of adding to 29.9 grams of sodium iodide dissolved in water 71 decigrams of chlorine?

$$2\ Na\,I + Cl_2 = 2\ Na\,Cl + 2\ I.$$

Ans.: 11.7 grams of sodium chloride, produced 25.3 grams of iodine set free.

50. To obtain 25.3 grams of iodine from sodium iodide, how much manganese dioxide, and how much sulphuric acid, should be used?

Ans.: 19.6 grams of $H_2\,SO_4$, 8.7 of $Mn\,O_2$.

51. How much calcium fluoride must be treated with sulphuric acid to obtain five grams of hydrofluoric acid?

$$Ca\,Fl_2 + H_2\,SO_4 = Ca\,SO_4 + 2\ H\,Fl.$$

Ans.: 9.75 grams.

52. How much red lead must be heated to obtain 3.2 grams of oxygen?

$$2\ Pb_3\,O_4\ (\text{heated}) = 6\ Pb\,O + O_2.$$

Ans.: 136.7 grams.

53. How much carbonic acid will be given off when a charge of 1000 pounds of carbonate of lime is heated in a lime-kiln, and how much quicklime will be produced?

$$Ca\,CO_3\text{ (heated)} = Ca\,O + CO_2.$$

Ans.: 560 lbs. CaO produced; 440 lbs. CO_2 given off.

54. If .325 gram of zinc be hung in a solution of lead nitrate, how much metallic lead will be substituted for the zinc?

$$Zn + Pb\,(NO_3)_2 = Pb + Zu\,(NO_3)_2.$$

Ans.: 1.0325 gram.

55. How much sodium hydrate and copper nitrate, respectively, must be used to obtain 1.94 gram of copper hydrate?

$$2\,NaHO + Cu\,(NO_3)_2 = Cu\,H_2\,O_2 + 2\,Na\,NO_3.$$

Ans.: 1.6. gram, Na HO. 3.74 Cu $(NO_3)_2$.

56. How much lead sulphate will be precipitated by the addition of sulphuric acid to a solution containing 1.322 gram of lead nitrate?

$$H_2\,SO_4 + Pb\,(NO_3)_2 = Pb\,SO_4 + 2\,H\,NO_3.$$

Ans.: 1.21 gram.

57. What is the per-centage composition of phosphoric acid?

(**Note.** Find the molecular weight of phosphoric acid, and the amount of phosphorus in the molecule: then state thus:

Mol. wt. phosphoric acid : amt. of P : : 100 : % of P. and so for each of the other elements present.)

Ans.: H, 3.06, &c. P, 31.63, &c. O, 65.31 nearly.

58. What is the per-centage composition of hydrochloric acid gas?

Ans.: Cl 97.26, &c. H 2.74 nearly.

59. A salt of copper has the following composition: Copper, 39.62 %; Sulphur, 20.125 %; Oxygen, 40.25 %. Find its formula.

(Note. Divide each per-centage number by the atomic weight of the element which it represents: the fractional numbers obtained are in the same ratio as the whole numbers expressing the number of atoms of each element to be set down in the formula of the salt. The simplest way to raise or reduce the fractional numbers to whole numbers is to divide each by the lowest. An error in the second or third place of decimals may be neglected as due to unavoidable errors of experiment.)

Ans.: $Cu\,SO_4$.

60. An acid of phosphorus has the following composition: Hydrogen, 3.66 %; Phosphorus, 37.80 %; Oxygen, 58.53 %. What is its symbol?

Ans.: $H_3\,PO_3$.

61. Find the symbol of a salt which has the following composition: Zinc, 40.37 %; Sulphur, 19.87 %; Oxygen, 39.75 %. Ans.: $Zn\,SO_4$.

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